ISIS

INDIVIDUALIZED SCIENCE INSTRUCTIONAL SYSTEM

It's All Air

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It's All Air

ANNOTATED TEACHER'S EDITION

Ginn and Company

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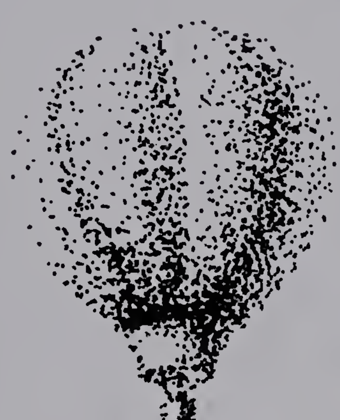
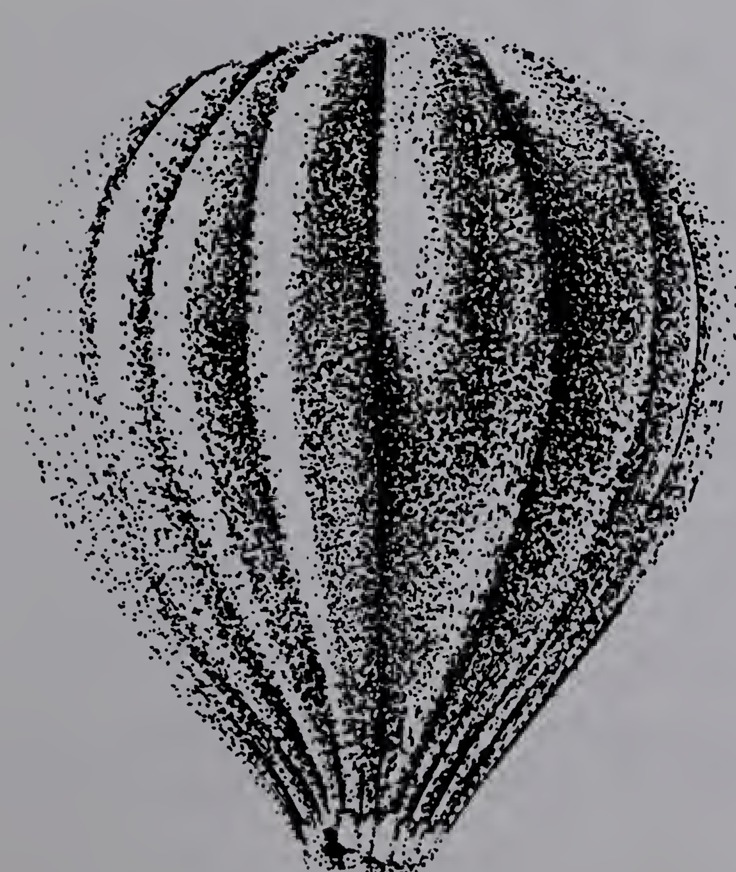
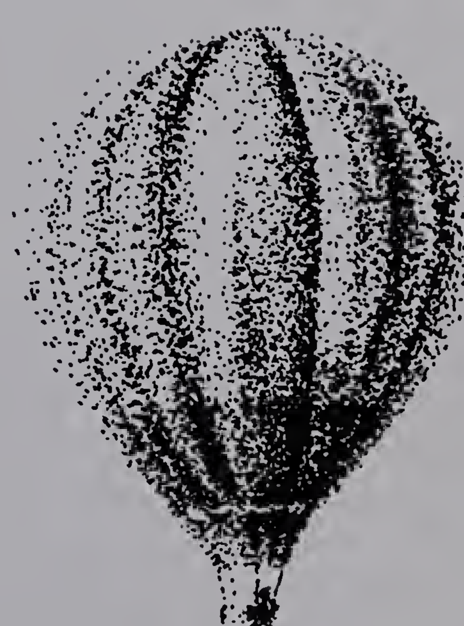






OVERVIEW

Most common gases are invisible, and students are often only vaguely aware that gases — particularly atmospheric gases — are vitally involved in their lives. Through investigations of a variety of different behaviors of the components of air, *It's All Air* attempts to provide students with a more concrete understanding of these substances they cannot see.

The filtering of solar radiation by the atmosphere, the air's role in the nitrogen and carbon dioxide–oxygen cycles, and air as a natural resource for chemicals and fertilizers are covered in *It's All Air*. The relation of air and its pollution to the standard of living is explored.

Air pollution, its effects, and its control are examined on as up-to-date a basis as possible. It is noted that though there are natural processes that can rid the air of most pollutants, the modern way of life has accelerated pollution emissions to a degree that makes these natural processes inadequate.

ORGANIZATION

It's All Air contains ten core activities, four advanced activities, and three excursion activities. The first activity in each section is a planning activity and should be done before any of the other activities in that section. Only in the advanced section is there a prescribed order. Students planning to do Activities 12 and 13 should do Activity 12 first.

About half of the core activities deal with the major components of air, particularly the characteristics and behaviors of these gases that relate to maintaining life. In the remainder of the core activities, the students investigate air pollutants.

In the advanced activities, pressure, volume, and temperature relationships are investigated, as well as the relative chemical reactivities of nitrogen, oxygen, and argon.

In the excursion activities, students can demonstrate differential air pressure by causing an empty can to collapse. And they can investigate the depletion of dissolved oxygen in water by oxygen-using materials.

MATERIALS AND EQUIPMENT

The following tables show the quantity and the frequency of use of each item used in each activity. The activities that require no materials are not listed in the tables.

It is important to collect and organize all the materials for each minicourse before the students begin any of the activities, since the students will be working simultaneously on different activities. Having all materials readily available allows students to do the activities in the order they choose. The amount of material you will need to make available will depend on the number of lab groups that will be doing each activity. As lab groups use the "skipping option" and as they scatter themselves

throughout the activities, the total amount of materials needed at one time for each activity will decrease.

CONSUMABLE ITEMS	MINIMUM MATERIALS PER LAB GROUP† PER ACTIVITY									
	2	7	8	9	10	12	13	14	16	17
Ammonia water, $NH_3\,(aq)$, concentrated (ml)			10							
Can, empty, thin aluminum beverage									1	
Candle, small	1									
Cotton plug for test tube		1								
Copper(II) sulfide, CuS (g)				1						
Dextrose (g)										1.5
Hydrogen peroxide, H_2O_2, 3% (ml)		5								
*Limewater (ml) made from:	100									
calcium hydroxide, $Ca(OH)_2$ (g)	0.5									
Litmus paper strip, blue				1						
Litmus paper strip, red								1		
Magnesium, granular (g)								2		
Magnesium ribbon (cm)								2		
Manganese dioxide, MnO_2 (g)		0.1								
Match, safety	1	2						1		
*Methylene blue solution in dropping bottle (ml)										1
Paper, graph (sheet)					1	1	1			
Paper (sheet)				1						
Potassium hydroxide, KOH (g)										2.5
Tape, masking, at least 2 cm wide (cm)									20	
Wood splint		1								

*See "Advance Preparations."
†A *lab group* is defined as one student, a pair of students, or any size group of students that you choose.

NONCONSUMABLE ITEMS	MINIMUM MATERIALS PER LAB GROUP* PER ACTIVITY										
	2	3	7	8	9	10	12	13	14	16	17
Balance					1				1		1
Barometer, classroom							1				
Bottle, wide mouthed	1										
Burner, Bunsen				1	1			1	1	1	
Clamp, ring, for ring stand									1	1	
Clamp, test tube, for ring stand								2			
Clay, about size of penny (piece)	1										
Clay triangle									1		
Copper wire, 18 gauge or heavier, clean (cm)				20							
Crucible with cover									1		
Filter, cyan		1									
Filter, magenta		1									
Filter, yellow		1									
Flask, Erlenmeyer, 250 ml				1							

*A *lab group* is defined as one student, a pair of students, or any size group of students that you choose.

NONCONSUMABLE ITEMS	MINIMUM MATERIALS PER LAB GROUP† PER ACTIVITY										
	2	3	7	8	9	10	12	13	14	16	17
Flask, Erlenmeyer, 500 ml											1
*Gas tube, special, made from the following:							1	1			
beaker, 600 ml							1				
mercury (ml)							0.1	0.1			
tubing, glass, 2-mm inside diameter (length in cm)							30	30			
Glove or mitt-type pot holder										1	
Graduated cylinder, 10 ml			1								
Graduated cylinder, 50 ml				1						1	
Graduated cylinder, 100 ml	1										1
Lid from can, metal									1		
Medicine dropper					1				1		
Pan, shallow	1										
Ring stand								1	1	1	
Ruler, metric							1	1			
Safety goggles	1		1	1	1		1	1	1	1	1
Scissors										1	
Spatula, steel, small									1		
Spectroscope		1									
Stirring rod, glass				1							
Stopper, rubber, one hole, to fit medium test tube					1						
Stopper, rubber, solid, to fit 500-ml flask											1
Test tube, 18 mm X 150 mm (medium)			2		1						
Test tube, 25 mm X 200 mm (large)								1			
Test-tube holder					1						
*Thermometer and special gas tube made from the following:								1			
gas tube, special (see above)								1			
rubber band, very small								2			
thermometer								1			
Tongs, beaker										1	
Tongs, crucible									1		
Triangle, 45°, 45°, 90°							1				
Tweezers									1		
Wire (cm)								40			
Wire gauze										1	
Resource Unit 2						1					
Resource Unit 4						1					
Resource Unit 5				1							
Resource Unit 10					1						
Resource Unit 17				1	1				1		
Actions and Reactions									1		

*See "Advance Preparations."

†A *lab group* is defined as one student, a pair of students, or any size group of students that you choose.

Activity 2

To make the limewater, put 5 g of calcium hydroxide, $Ca(OH)_2$, into a 1000-ml narrow-mouthed container. Add tap water to make one litre of solution. Stopper the container, and shake it vigorously. Let the container stand overnight. Any undissolved solids will settle to the bottom. Carefully decant the clear limewater into dispensing bottles.

Activities 12 and 13

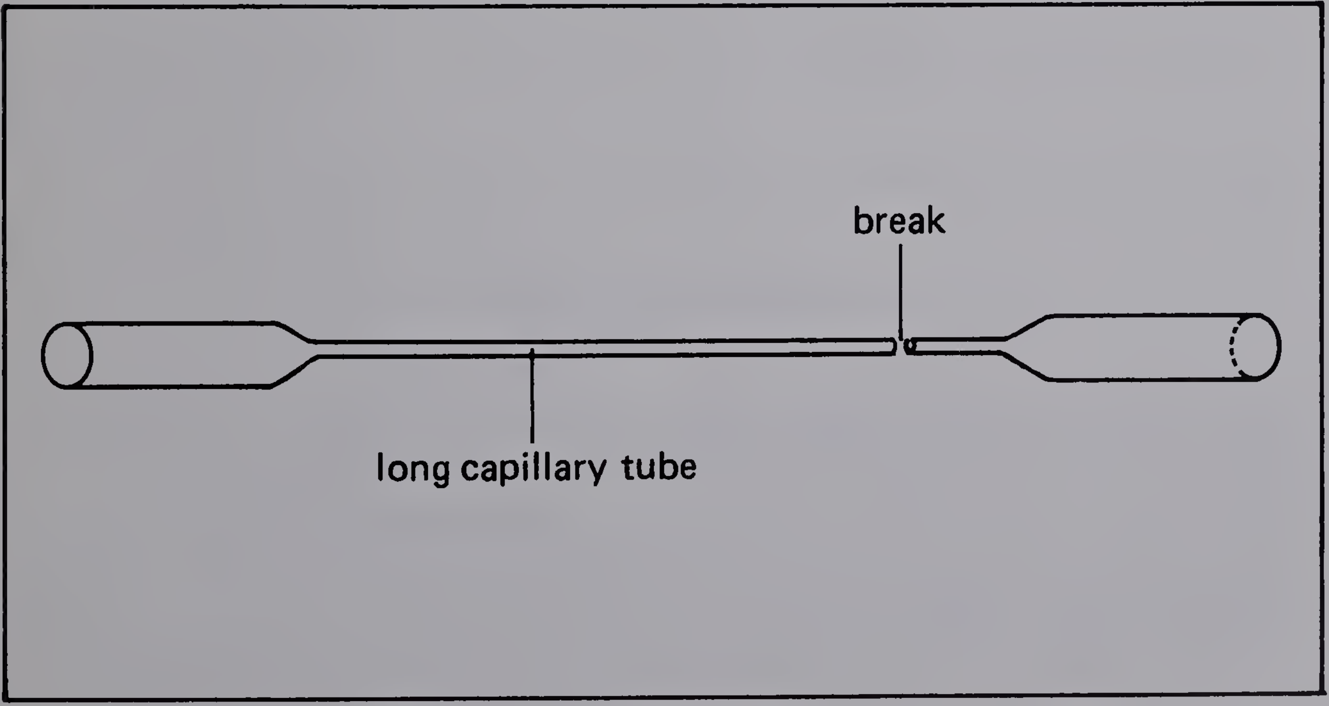
Both in your advance preparations and in students' investigations, the goal is to minimize exposure hazards and mishandling of metallic mercury. Mercury can be handled safely. See the special procedures for handling mercury spills on page ATE 7 in "Background Information."

Select for Activities 12 and 13 the laboratory site that will minimize hazards. Supervise closely students known to be careless or those who might be tempted by the attractiveness of the mercury.

All glassware and the mercury must be kept scrupulously clean.

To make gas tubes, cut 30-cm lengths of 2-mm inside diameter (or 3-mm outside diameter with 0.75 mm walls) glass tubing. Seal one end by melting it.

Make a dropping pipette for filling the gas tubes by heating a 15-cm piece of 3 to 4 mm tubing until it softens. Remove the tube from the flame and immediately pull it, preferably vertically, to form a long capillary. Notch it with a file near one end and break it.



Fit a medicine dropper bulb on the large end, and very slowly fill the capillary with mercury. (A rubber-bulb ear syringe with a fine or properly tapered end can also be used.) Insert the end of the capillary one-third of the way down the prepared gas tube (sealed and tested for closure at one end), keeping both as nearly horizontal as possible. Discharge mercury from the capillary so that the mercury plug measures 2 cm to 3 cm in length.

During the filling, if there is separation or any mercury falls to the bottom of the gas tube, the capillary can be used without the bulb to introduce air past the plug. Insert the capillary all the way to the bottom of the gas tube, tilt the tube slightly so that the mercury flows toward the open end, and remove the capillary when the fallen mercury has been deposited with the plug 15 to 20 cm from the sealed end of the gas tube. A number 20 copper wire, preferably lacquered against amalgamation, can also be used to gather together separated mercury plugs.

Gas tubes should be dispensed sealed end down in the 600-ml beaker that students should be instructed to use as a spill saver.

For Activity 13, the thermometer and gas tube assembly should be tied together by two very small rubber bands as illustrated on page 60. The thermometer bulb should be about halfway up the column of air trapped by the mercury plug.

Activity 17

Dissolve one gram of methylene blue powder in enough water to make 100 ml of solution. Dispense it in a dropping bottle.

BACKGROUND INFORMATION

Safety Contracts

ISIS strongly recommends that all students sign a safety contract early in the school year before beginning any minicourse. A sample safety contract can be found in *Managing ISIS*.

To help enforce the provisions of the safety contract, ISIS has placed various cautionary notes in the student's book. You will note that these cautions emphasize eye safety in particular.

Eye Safety

ISIS recommends that students wear safety goggles whenever they are working in a laboratory situation. Although a particular student may not be working with hazardous materials, other students nearby may be.

In Activities 3 and 14, students have to be warned that illumination that is very bright and actinic (containing ultraviolet rays) should not be viewed directly. Looking at the sun or burning magnesium can damage eyes.

Working with Chemicals

Early in the school year, spend some time instructing your students on general laboratory safety and on appropriate precautions for working safely with chemicals. There are several general safety suggestions in *Managing ISIS*. Specific suggestions for this minicourse follow.

In Activities 12 and 13, students work with a gas tube containing a plug of mercury. If the tube is handled carefully, there should be no problem of mercury spills. However, if a spill should occur, follow this procedure, since mercury is extremely toxic.

1. Draw larger droplets of the mercury into a medicine dropper or a suction device. (Instructions for making a suction device are given below.)
2. Then liberally dust the area of the spill with powdered sulfur or zinc dust. Work the dust into all cracks and crevices. Leave it for twenty-four hours.
3. Sweep up the sulfur (and mercury sulfide or zinc amalgam), and dispose of it as solid waste.

A suction device can be made with a wash bottle and a faucet aspirator or vacuum pump. Attach the suction tube to the blowing tube of a glass wash bottle.

Be alert to the temptation some students may have to handle mercury. It is extremely toxic. It also ruins metal jewelry by coating it and weakening it.

Disposal and Conservation of Materials

You will have to direct students on methods for safely conserving and disposing of various liquids and solids. Refer to *Managing ISIS* for general suggestions. Specific suggestions for this minicourse follow.

Mercury is easily conserved and should be recycled. It can be cleaned of surface scum by a wet run followed by a dry run through a pinhole in filter paper in a funnel. Amalgamated metals and persistent dirt can be removed by storing pooled mercury in a beaker under a small amount of dilute nitric acid in a hood. After neutralization, the pinhole technique should work. Wear safety goggles during this procedure.

EVALUATION SUGGESTIONS

In addition to the *Minicourse Test*, answers to which are provided with the test, you may want to use the following essay questions and laboratory performance item.

Essay Questions

Five essay questions are included here with model answers for your convenience. The first four relate to core activities. The fifth relates to an advanced activity.

1. The air traps certain kinds of energy. What might be the effects on living things, and why, if the amount of ozone in the air was decreased?

Answer: Reduction in the amount of ozone would allow more ultraviolet light to reach Earth. The result could be severe sunburn and increased occurrence of skin cancer.

2. Nitrogen is the fourth most abundant element in the human body. It is a necessary part of human protein. Yet, human beings cannot use the nitrogen gas in the air. Why can't they? How do human beings get the nitrogen they need?

Answer: The nitrogen in the air is in the form of tightly bound pairs of atoms, which are fairly unreactive. Human beings and other animals cannot break these pairs of atoms (diatomic molecules) apart. Instead, human beings must get the essential nitrogen by eating plants that contain nitrogen compounds or by eating other animals that in turn had eaten plants.

3. Since people know that pollutants in the air are harmful to living things, why haven't they eliminated all sources of pollutants to get clean air?

Answer: Many pollutants are by-products of manufacturing and energy-producing processes and of other burning processes, such as fuels burning in cars. Eliminating all these sources of pollution would mean doing without many things, not only things that are nice to have, but things that are essential to health and well-being, such as hospital services, adequate food, heat in cold weather, and so on. People have to decide which things are worth some pollution, which are not, and what degree of pollution control they can afford, both in terms of money and of energy.

4. Describe at least three processes that are important in cycling oxygen and carbon dioxide to and from the air.

Answer: In sunlight, plants produce oxygen and take in carbon dioxide. Plants also produce carbon dioxide and take in oxygen. Animals produce carbon dioxide and take in oxygen. Decay microorganisms produce carbon dioxide and take in oxygen. The burning of fossil fuels produces carbon dioxide and uses oxygen.

5. When heated in air, many metals react with the oxygen, but few react with the nitrogen. Explain this behavior in terms of the structures of oxygen and nitrogen.

Answer: The nitrogen in the air, in the form of diatomic molecules, is relatively unreactive because it has a strong triple bond holding the two atoms together. This triple bond is a result of the three unpaired electrons of each nitrogen atom pairing with the three unpaired electrons of another nitrogen atom. Oxygen, also occurring in air as diatomic molecules, is more reactive because the bond between the atoms in the diatomic molecule is much weaker and also because there is an unpaired electron on each of the atoms in the molecule.

Laboratory Performance

The laboratory performance item requires individual students to demonstrate skills used in the minicourse. You may wish to assess students soon after they have learned the skills.

Provide students with graph paper. Ask them to construct a graph of a set of data, such as the sample data given below, and then interpret the graph in terms of a trend.

U.S. AIR POLLUTION TRENDS (in billions of kg of pollutant released)		
Year	SO_x	Particulates
1940	20	25
1950	22	24
1960	21	23
1970	31	23

REFERENCES

Andrews, William A., ed. *A Guide to the Study of Environmental Pollution.* Englewood Cliffs, New Jersey: Prentice-Hall, Inc., 1972.

Chapter 4 is a student's and teacher's reference. It includes most aspects of pollution dealt with in the minicourse and is accompanied by field and laboratory studies.

Beckmann, Peter, ed. *The Health Hazard of NOT Going Nuclear.* Boulder, Colorado: Golden Press, 1976.

This book stresses that increased burning of coal in the United States may be more dangerous to human health than an alternative use of nuclear-reactor power.

Biological Science Curriculum Study. *Biological Science: An Ecological Approach.* Chicago. Rand McNally & Co., 1978.

A complete student reference on the biological importance of the gas cycles is included.

Bolin, Bert. "The Carbon Cycle." *Scientific American* 222 (September 1970): 124-32.

This reference for teachers is also available in *Chemistry in the Environment, Readings from Scientific American*, published in 1973 by W.H. Freeman and Company.

Council on Environmental Quality. *Annual Report of the Council on Environmental Quality.* Washington, D.C.: Government Printing Office.

This annual report on the state of the environment gives both regional and national information on pollution trends. It contains both general information and specific data on air and other environmental pollutants.

Delwiche, C.C. "The Nitrogen Cycle." *Scientific American* 222 (September 1970): 137-46.

This reference for teachers and able students is also available in *Chemistry in the Environment; Readings from Scientific American*, published in 1973 by W.H. Freeman and Company.

Eubanks, I. Dwaine, and Dermer, Otis C. *Chemistry in Civilization*. New York: Ronald Press, 1975.

This is an introductory text for students with no background in science. It treats most of the topics covered in this minicourse.

Garells, Robert M.; MacKenzie, Fred T.; and Hunt, Cynthia. *Chemical Cycles and the Global Environment.* Los Altos, California: William Kaufmann, Inc., 1975.

This is a development of biogeochemical data and problems on a multidisciplinary basis.

Kormondy, E.J. *Concepts of Ecology.* Englewood Cliffs, New Jersey: Prentice-Hall, 1969.

Pages 40 through 48 include a discussion of the gaseous nutrient cycles. It is an appropriate reference for teachers.

Manufacturing Chemists Association. *Air Pollution: Causes and Cures.* Manufacturing Chemists Association, Washington, D.C., 1973.

This is a concise and well-illustrated reference appropriate for students.

E.Q. Index. Washington, D.C.: National Wildlife Federation.
The *E.Q. Index* has been published yearly since 1969, each year evaluating the progress in improving environmental quality. Students should find this interpretation of environmental quality trends interesting. A single copy can be obtained free from Educational Services, National Wildlife Federation, 1412 16th Street, N.W., Washington, DC 20036.

Pryde, Lucy T. *Environmental Chemistry: An Introduction.* Menlo Park, California: Cummings Publishing Co., 1973.
This is a comprehensive reference for able students. The section on air is also available as a separate supplement called *Chemistry of the Air Environment.*

U.S. Environmental Protection Agency. *Progress in the Prevention and Control of Air Pollution.* Washington, D.C.: Government Printing Office.
These annual reports made to Congress by the EPA outline progress in pollution research and control. Each pollution source is evaluated with respect to compliance with federal regulations. The report also summarizes state and local control programs.

Weaver, Elbert C. *Environmental Pollution.* New York: Holt, Rinehart, and Winston, 1971.
This is an elementary, activity-centered introduction to environmental pollution. Pages 1 through 40 deal with air pollution.

The U.S. Environmental Protection Agency publishes a number of useful reports and brochures on air pollution. An annotated bibliography is available from them. EPA publications can be obtained by writing to the U.S. Environmental Protection Agency, Office of Public Affairs, Washington, DC 20460. The following are some of their publications on the topic.

Clean Air and Your Car.
This discusses auto pollutants and their effects, emission control systems and their costs in gas mileage, fuel, and sticker price, and strategies an average driver can use to reduce fuel consumption and produce cleaner air.

Clean Air: The Breath of Life.

This presents a short summary of the nature of air pollutants, current emissions, and their relationship to federal regulations.

Health Effects of Air Pollutants.

This presents a concise discussion of the effects of the major air pollutants on human beings.

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In addition to the major effort by the ISIS permanent staff, writing conference participants, and author-consultants (listed on the inside of the back cover), the following contributed to this minicourse.

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FOREWORD

Evidence has been mounting that something is missing from secondary science teaching. More and more, students are rejecting science courses and turning to subjects that they consider to be more practical or significant. Numerous high school science teachers have concluded that what they are now teaching is appropriate for only a limited number of their students.

As their concern has mounted, many science teachers have tried to find instructional materials that encompass more appropriate content and that allow them to work individually with students who have different needs and talents. For the most part, this search has been frustrating because presently such materials are difficult, if not impossible, to find.

The Individualized Science Instructional System (ISIS) project was organized to produce an alternative for those teachers who are dissatisfied with current secondary science textbooks. Consequently, the content of the ISIS materials is unconventional as is the individualized teaching method that is built into them. In contrast with many current science texts which aim to "cover science," ISIS has tried to be selective and to limit our coverage to the topics that we judge will be most useful to today's students.

Obviously the needs and problems of individual schools and students vary widely. To accommodate the differences, ISIS decided against producing tightly structured, pre-sequenced textbooks. Instead, we are generating short, self-contained modules that cover a wide range of topics. The modules can be clustered into many types of courses, and we hope that teachers and administrators will utilize this flexibility to tailor-make curricula that are responsive to local needs and conditions.

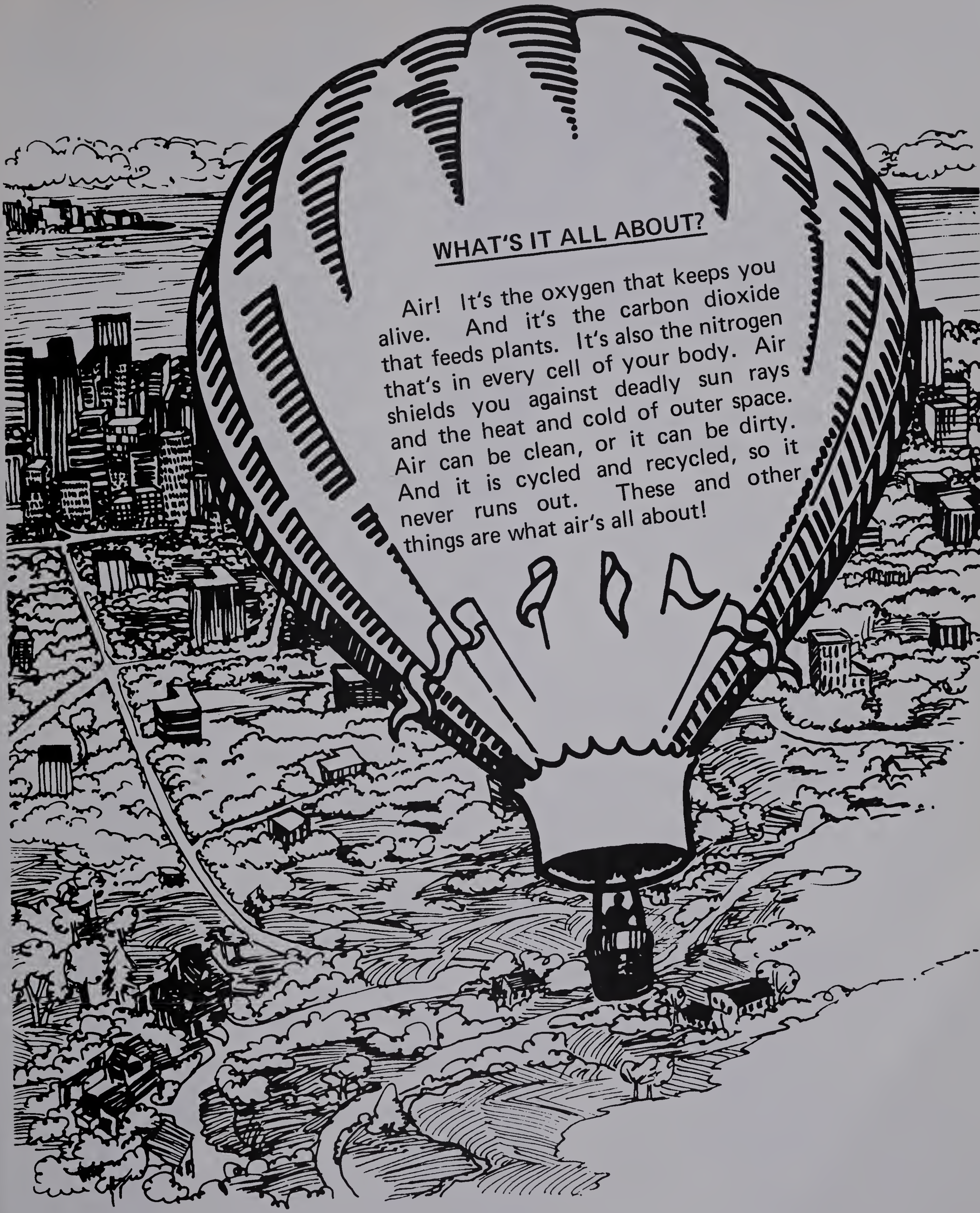
ISIS is a cooperative effort involving many individuals and agencies. More than 75 scientists and educators have helped to generate the materials, and hundreds of teachers and thousands of students have been involved in the project's nationwide testing program. All of the ISIS endeavors have been supported by generous grants from the National Science Foundation. We hope that ISIS users will conclude that these large investments of time, money, and effort have been worthwhile.

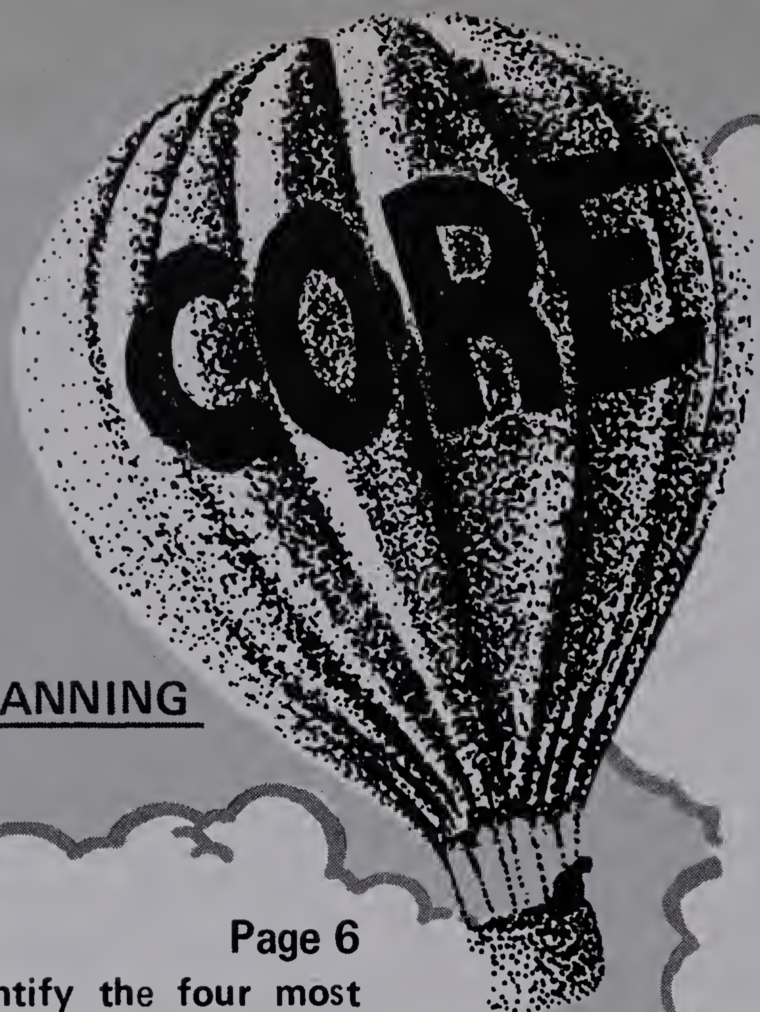
Ernest Burkman
ISIS Project
Tallahassee, Florida

CONTENTS PAGE

WHAT'S IT ALL ABOUT?

Air! It's the oxygen that keeps you alive. And it's the carbon dioxide that feeds plants. It's also the nitrogen that's in every cell of your body. Air shields you against deadly sun rays and the heat and cold of outer space. Air can be clean, or it can be dirty. And it is cycled and recycled, so it never runs out. These and other things are what air's all about!





ACTIVITY 1: PLANNING

Activity 2 **Page 6**

Objective 2-1: **Identify the four most plentiful gases in dry air and rank them according to their abundance.**

Sample Question: List the four major gases in dry air in order of their abundance from greatest to least.

Objective 2-2: **Describe the chemical reactivity of the main components of dry air.**

Sample Question: Match each gas with its reacting behavior in dry air.

Gas	Behavior
A. Argon	1. reacts readily
B. Carbon dioxide	2. does not react readily
C. Oxygen	

Activity 3 **Page 9**

Objective 3-1: **Tell which kinds of solar radiation are filtered out by the air and what the effects on living things would be if such filtering did not occur.**

Sample Question: Choose the type of radiation that is mostly filtered out by the gases in air. Then match this type with the consequences to Earth if filtering did not occur.

Type of Radiation	Consequences If *Not* Filtered
A. Ultraviolet	1. The temperature on Earth would rise, killing living things.
B. Visible light	2. People would suffer severe sunburn and more skin cancer.
C. Infrared	3. Living things would not get the oxygen they need and would suffocate.
	4. The temperature on Earth would drop, freezing living things.

Objective 3-2: **Explain why it was difficult to predict that the use of aerosol sprays lessens the air's ability to keep harmful solar radiation from damaging life.**

Sample Question: It was difficult to predict that the use of chlorofluorocarbon aerosol sprays would increase skin cancer. This was because

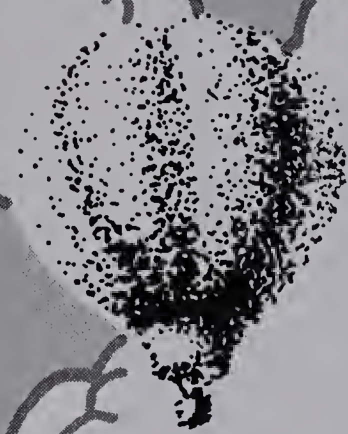
A. chlorofluorocarbons are invisible to the naked eye.
B. scientists don't usually have to make predictions.
C. chlorofluorocarbons decompose into harmless substances.
D. scientists didn't know about some of the chlorofluorocarbon reactions that could take place in the environment.

Activity 4 — Page 14

Objective 4-1: Describe methods for detecting solid and gaseous air pollutants.

Sample Question: What is an accurate method of detecting gaseous air pollutants?

A. Use an electrostatic precipitator.
B. Check for irritation of eyes and throat.
C. Note light-absorption patterns.
D. Use sticky tape to collect and measure them.



Activity 5 — Page 19

Objective 5-1: Tell how to reduce the amount of each of the five major air pollutants that get put into the air.

Sample Question: Match each air pollutant with a method of reducing it.

Pollutant	Remedy
A. Particulates	1. Burn fuel at high temperatures in plenty of air.
B. Hydrocarbons	2. Use an electrostatic precipitator.
C. Nitrogen oxides	3. Use a catalytic converter.
	4. Remove sulfur before burning the fuel.

Activity 6 — Page 25

Objective 6-1: Describe the effects of air pollutants (sulfur dioxide, nitrogen oxides, carbon monoxide, particulates, and hydrocarbons) on plants and animals.

Sample Question: How can sulfur dioxide cause damage to human beings?

A. It reduces the oxygen-carrying capacity of blood.
B. It irritates and damages the respiratory system (especially the nose, throat, and lungs).

Objective 6-2: Describe the effects of smog on people and plants.

Sample Question: Tell whether the following statement is true or false. Certain kinds of smog form acid rains that wash into the soil and damage plants.

Answers: 2-1. Nitrogen, oxygen, argon, carbon dioxide; **2-2.** A2, B1, C1; **3-1.** A2, **3-2.** D; **4-1.** C; **5-1.** A2, B1, C3; **6-1.** B; **6-2.** True

Activity 7 Page 28

Objective 7-1: Describe the oxygen–carbon dioxide cycle and its importance.

Sample Question: Match each process with its role in the oxygen–carbon dioxide cycle.

Process	Role in O_2–CO_2 Cycle
A. Respiration	1. Carbon dioxide is taken from the atmosphere and oxygen is added.
B. Photosynthesis	2. Oxygen is taken from the atmosphere and carbon dioxide is added.
C. Burning	
D. Decay after death	

Objective 7-2: Identify important practical uses of oxygen and carbon dioxide.

Sample Question: Match each substance with one or more of its uses.

Substance	Use
A. Oxygen	1. refrigeration
B. Carbon dioxide	2. medical treatment
	3. producing hot flames
	4. extinguishing fires

Activity 8 Page 33

Objective 8-1: Describe the processes in the nitrogen cycle and the importance of the cycle to plants and animals.

Sample Question: In the nitrogen-fixation process,

A. nitrogen in the air is converted into nitrogen compounds that plants can use.
B. nitrogen compounds in plants are converted into free nitrogen by lightning.
C. bacteria in plants fix the broken parts of nitrogen compounds in the roots.
D. the oxygen in the air is used up so that the nitrogen is left free for plants.

Objective 8-2: Describe, in terms of atoms, how nitrogen from the air is converted into a form useful to plants.

Sample Question: The nitrogen in the air can be used by plants

A. as soon as it reaches their leaves.
B. only after it is converted into tightly bound pairs of atoms, N_2.
C. only after the tightly bound pairs of N atoms in the air are separated.
D. only after it is separated from the oxygen in the air.

Objective 8-3: Name four uses of nitrogen or nitrogen compounds.

Sample Question: Which of the following are uses of nitrogen gas or nitrogen compounds?

A. Explosives
B. Fertilizer
C. Refrigeration
D. Animal respiration

Activity 9 **Page 39**

Objective 9-1: Describe how the production of things people need or want relates to air pollution.

Sample Question: People's desires for better and easier living have

A. decreased air pollution.
B. increased air pollution seriously.
C. not affected air pollution.
D. increased air pollution, but not seriously.

Objective 9-2: Describe several factors that should be considered in making decisions to control air-pollution levels.

Sample Question: Suppose that a clothing factory in your town is releasing polluting gases into the air. What should the town government do?

A. It should close down the factory.
B. No action is required since wind carries the gases away.
C. It should require the factory to install expensive pollution-control devices.
D. None of the above are completely satisfactory.

Activity 10 **Page 45**

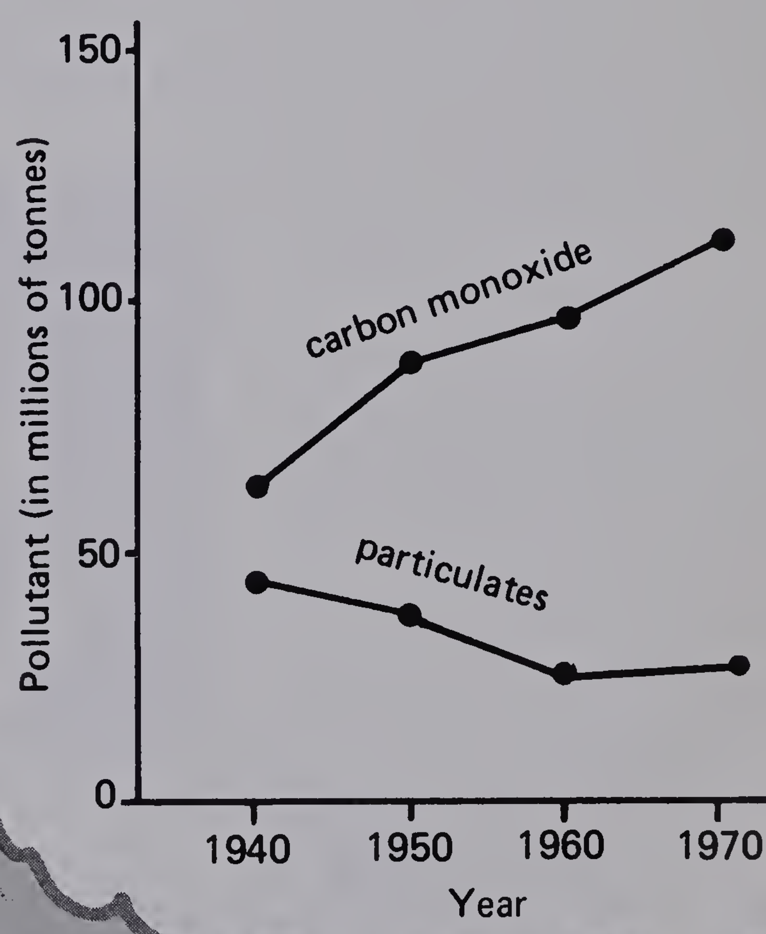
Objective 10-1: Describe trends in the amounts of five major air pollutants released in the United States in recent years.

Sample Question: Match the air pollutant with the trend in the amount of the pollutant since 1970.

Pollutant	Trend Since 1970
A. Sulfur oxides	1. amount increased
B. Hydrocarbons	2. amount decreased
	3. amount stayed the same

Objective 10-2: Using information in a graph, decide which air pollutants are being most successfully controlled in the United States.

Sample Question: Look at the graph below. Between 1940 and 1970, which pollutant was more successfully controlled?



Answers: 7-1. A2, B1, C2, D2; **7-2.** A2, 3, B1, 4; **8-1.** A; **8-2.** C; **8-3.** A, B, C; **9-1.** B; **9-2.** D; **10-1.** A2, B2; **10-2.** Particulates

ACTIVITY EMPHASIS: Air is composed of a relatively constant mixture of gases. The four most abundant are nitrogen, oxygen, argon, and carbon dioxide. Oxygen and carbon dioxide react readily; nitrogen and argon do not.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter. See "Advance Preparations" in ATE front matter.

Check the limewater reagent bottle periodically. If the limewater gets cloudy, replace it with a fresh supply.

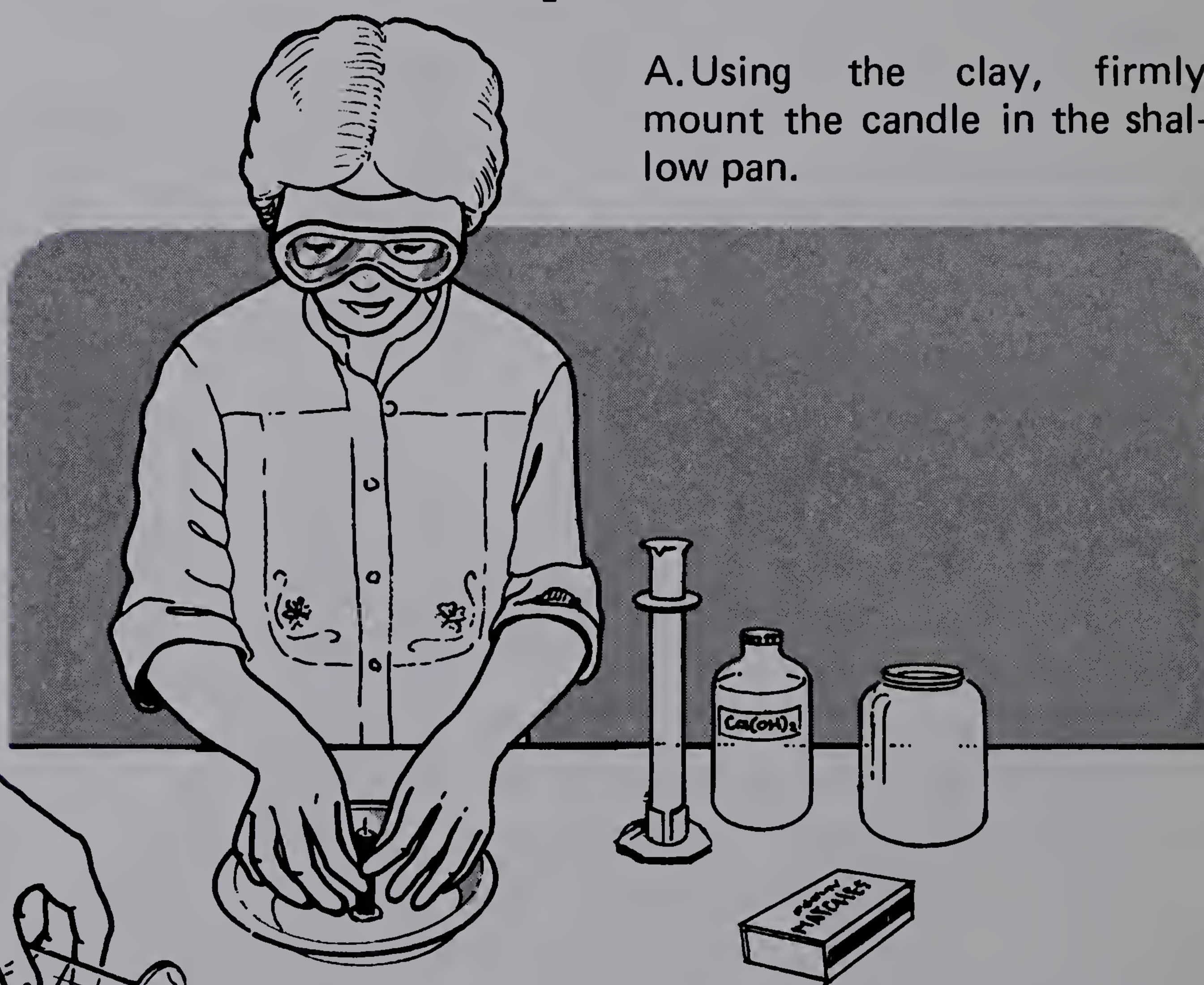
ACTIVITY 2: CONTENTS ALOFT

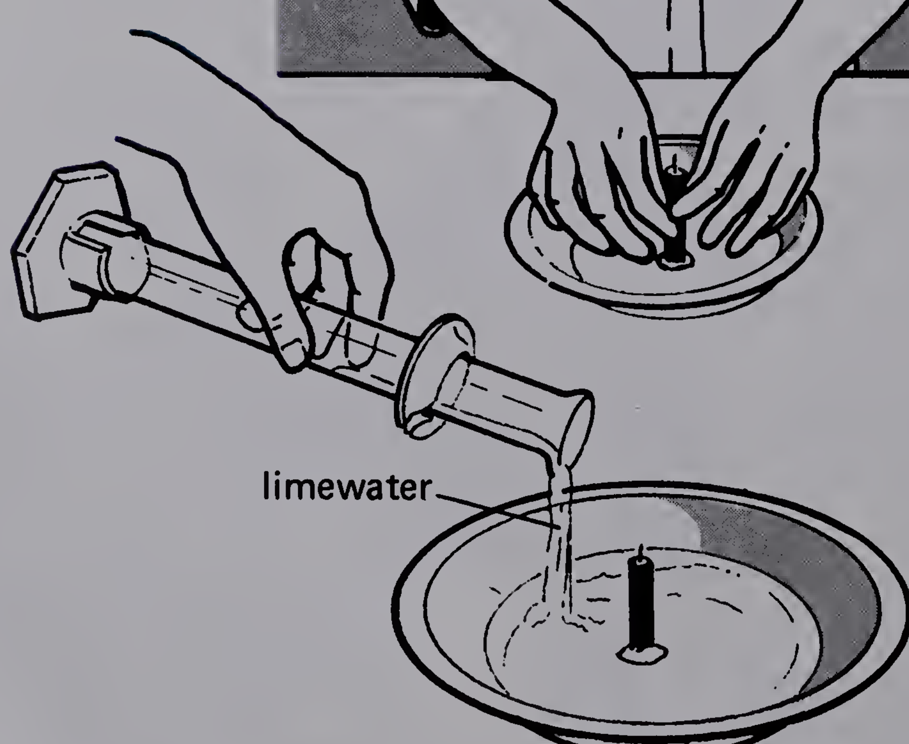
You live at the bottom of a big ocean of air. You can't see air. But if you fan it past your ear, you can hear it. You can feel it when the wind blows. And you can see a kite, a bird, or an airplane held up by air.

Think about what air is made of — one thing or many. You can find out. You'll need these materials.

safety goggles
100-ml graduated cylinder
wide-mouthed bottle at least twice as tall as the candle
shallow pan
safety matches
piece of clay about the size of a penny
small candle
100 ml limewater, $Ca(OH)_2$

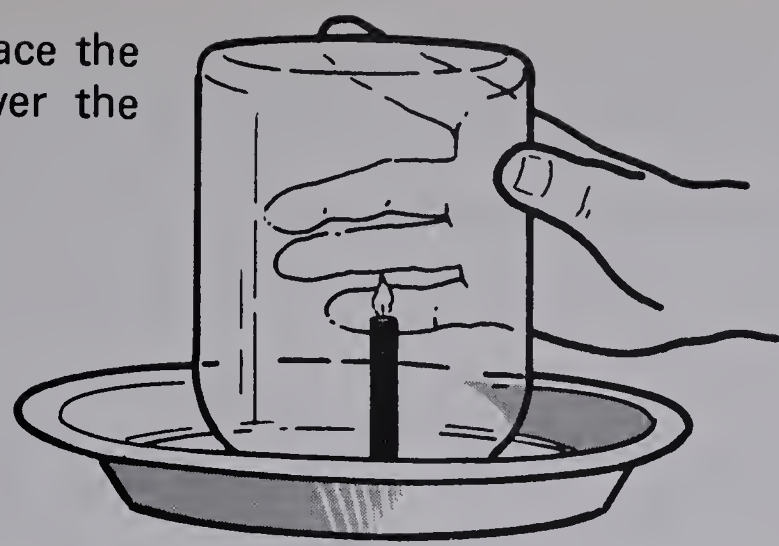
A. Using the clay, firmly mount the candle in the shallow pan.





B. Pour the 100 ml of limewater into the shallow pan.

C. Light the candle. Place the bottle upside down over the candle.



● 2-1. What happens to the water level at the mouth of the bottle as the candle burns?

2-1. It rises.

The products of the reaction are CO_2 and H_2O.

● 2-2. What happens to the limewater? To the candle flame?

2-2. The limewater gets cloudy. The candle flame goes out.

D. After the candle flame goes out, gently slide the bottle back and forth without lifting it out of the limewater.

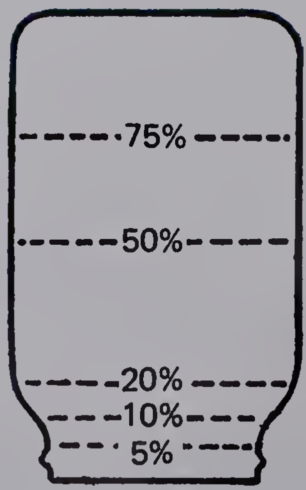
The cloudiness is suspended calcium carbonate precipitate: $Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O$.

● 2-3. What does the limewater look like now?

2-3. Very cloudy

The limewater rose in the bottle as the candle burned. Thus, there must have been less and less air above the limewater in the bottle. One way of explaining the decrease in air in the bottle is to assume that the burning candle used up the air.

But then why did the candle flame go out before all the air was used up? Perhaps air contains more than one thing. One thing is used up in burning, but the other is not.



● 2-4. Estimate what part of the air in the bottle — 5%, 10%, 20%, 50%, or 75% — was used up by the burning candle.

2-4. About 10% to 20%

The part of the air that was used up as the candle burned is called *oxygen*, O_2. By doing experiments that picked out one or the other of the gases in air, as you just did, scientists determined that air is a mixture of several gases.

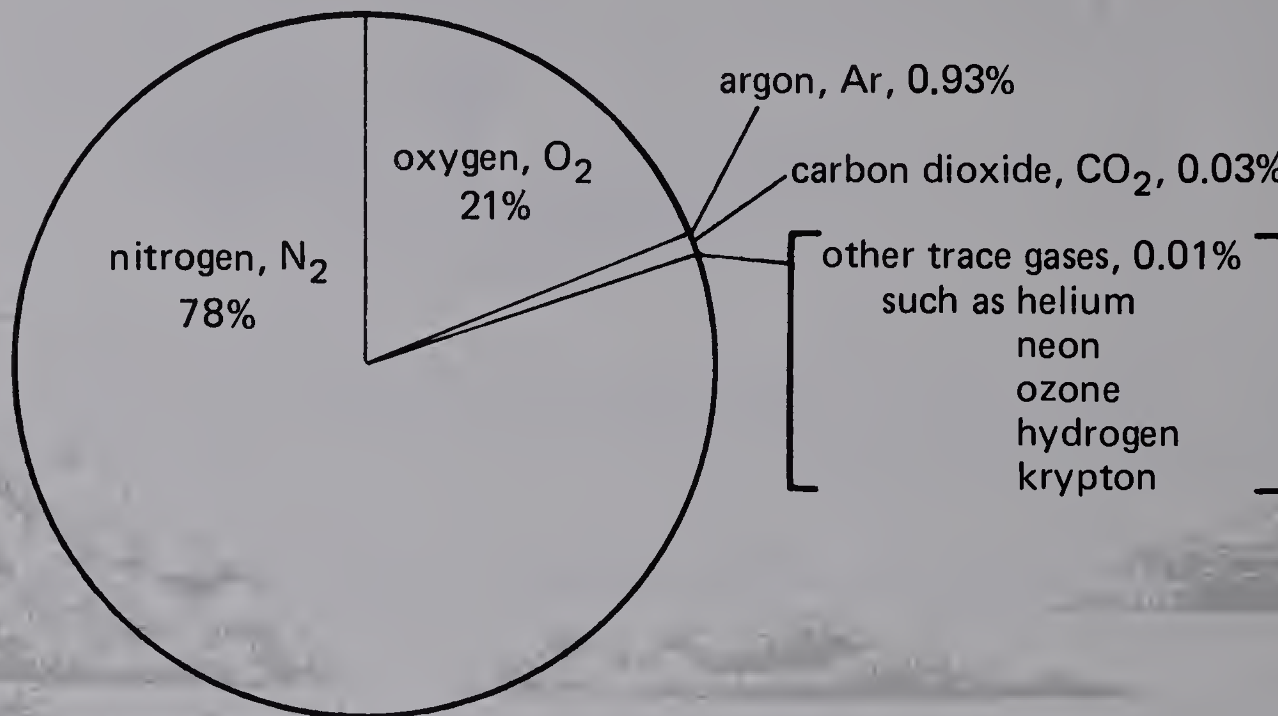


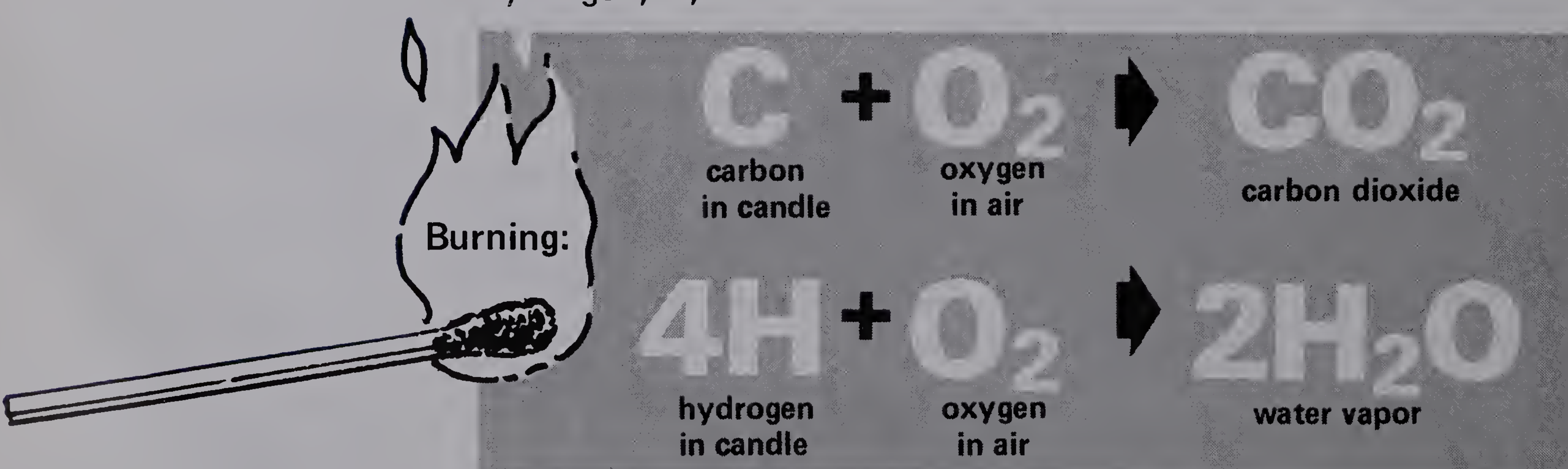
Figure 2-1

Figure 2-1 above shows the percentage by volume of gases in *dry* air at sea level. There is also water vapor in most air. But its percentage varies, as you can tell by changes in humidity. In wet air, there is often more water vapor than carbon dioxide — and sometimes even more than argon.

2-5. Nitrogen, oxygen, argon, and carbon dioxide

☆**2-5. List the four major gases in dry air in order from the most abundant to the least abundant.**

Now look back at your investigation. What the oxygen, O_2, reacted with in the burning candle was the carbon, C, and the hydrogen, H, in the candle and the wick.

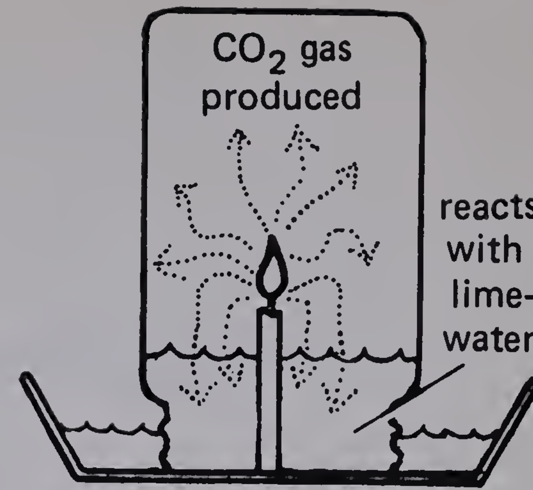


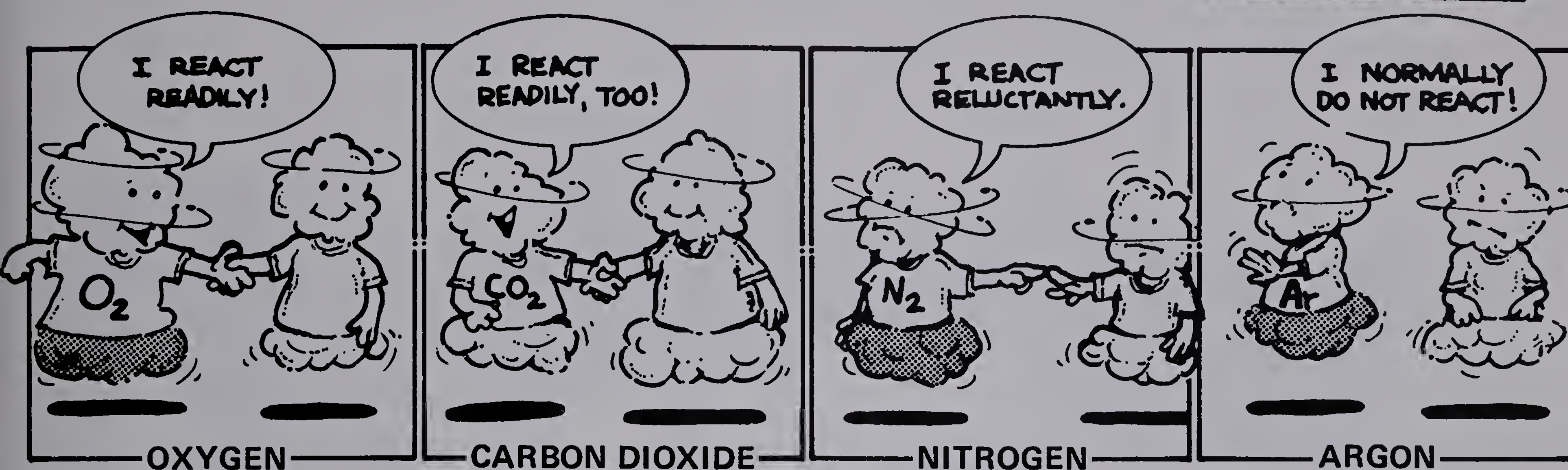
2-6. Carbon dioxide, CO_2, and water vapor, H_2O

● 2-6. When the candle reacted with oxygen (burned), what two substances were formed?

The carbon dioxide that was formed reacted with the limewater. The cloudiness of the limewater was evidence of this reaction.

So you've seen that two of the most abundant gases in air react readily — oxygen in the burning of the candle and carbon dioxide in the clouding of limewater.





The nitrogen inside the bottle was not changed by either the burning candle or the limewater. This behavior is not unusual for nitrogen. It does not react readily.

Of the four major gases in air, that leaves argon. Argon doesn't normally react with anything. And some of the other trace gases in air, such as helium and neon, don't react either.

The reluctance of nitrogen to react is due in part to the fact that N_2 molecules are joined by a triple bond. The reluctance of argon to react is due to the stability of its completed outer electron shell.

☆ **2-7. Match each gas with its reacting behavior.**

Gas	**Behavior**
A. Nitrogen	**1. reacts readily**
B. Oxygen	**2. does not react readily**
C. Carbon dioxide	
D. Argon	

2-7. A2, B1, C1, D2

ACTIVITY 3: AIR TO THE RESCUE

ACTIVITY EMPHASIS: Air keeps deadly sun rays from destroying life on Earth. How this protection might be destroyed is discussed.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter.

The sun sends a huge amount of energy to Earth, so much that if it all reached Earth, people would burn up. But air to the rescue!

Air traps some of the sun's energy so that it doesn't reach Earth. The rest strikes Earth, providing the heat and light necessary for life.

The sun's energy takes many different forms. Visible light is the only part that people can see. Look at Figure 3-1 below.

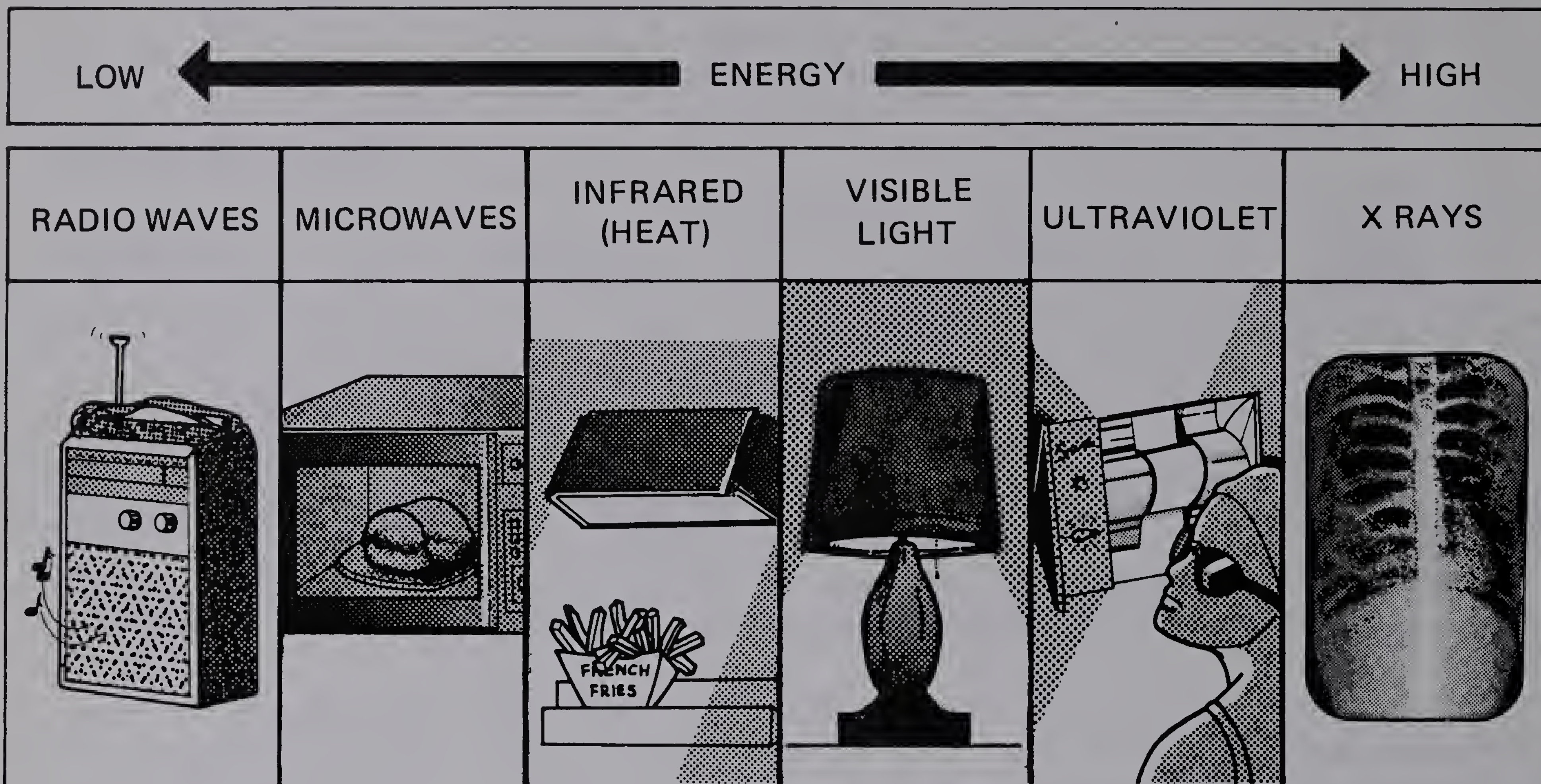


Figure 3-1

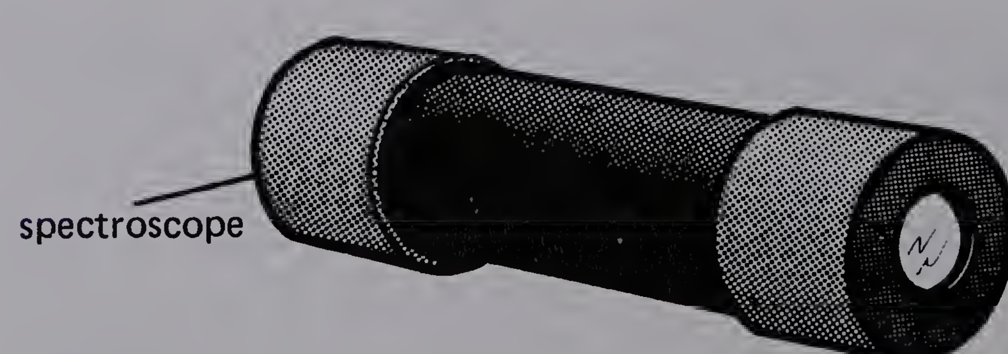
X rays and high-energy ultraviolet light are ionizing radiation with enough energy to destroy covalent bonds. Lower-energy ultraviolet and visible light excite electrons in chemical compounds but do not normally break chemical bonds. Infrared light causes increased vibration and rotation of bound atoms.

The sun need not be shining, but if it is very dark, students may need to use classroom lights as sources. An incandescent light is best, as fluorescent will not only give the full spectrum, but mercury emission lines as well (bright bands in the blue, green, and yellow).

To understand how air can trap (filter) some of the sun's energy, you can investigate visible light. You can read this page because visible light isn't filtered by air. Visible light, sometimes called *white light*, is really a mixture of several different colors. You can split white light into those separate colors. And then you can see how solar radiation can be filtered. You'll need these materials.

spectroscope
cyan (greenish blue) filter
magenta (reddish purple) filter
yellow filter

CAUTION	Never look directly at the sun, even through a spectroscope.



A. Point the spectroscope at a window, but not at the sun. Look through the end with the round opening.

• 3-1. When you look through the spectroscope, what do you see on either side of the slit?

3-1. Bands of colors (violet, blue, green, yellow, orange, and red)

The colors you saw make up the color spectrum of visible light. Look at Figure 3-2 below. But concentrate on the three wide areas of blue, green, and red.

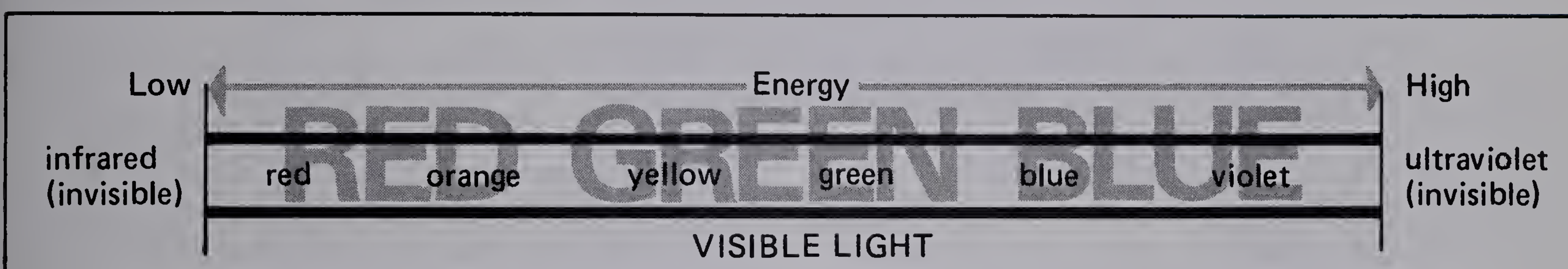
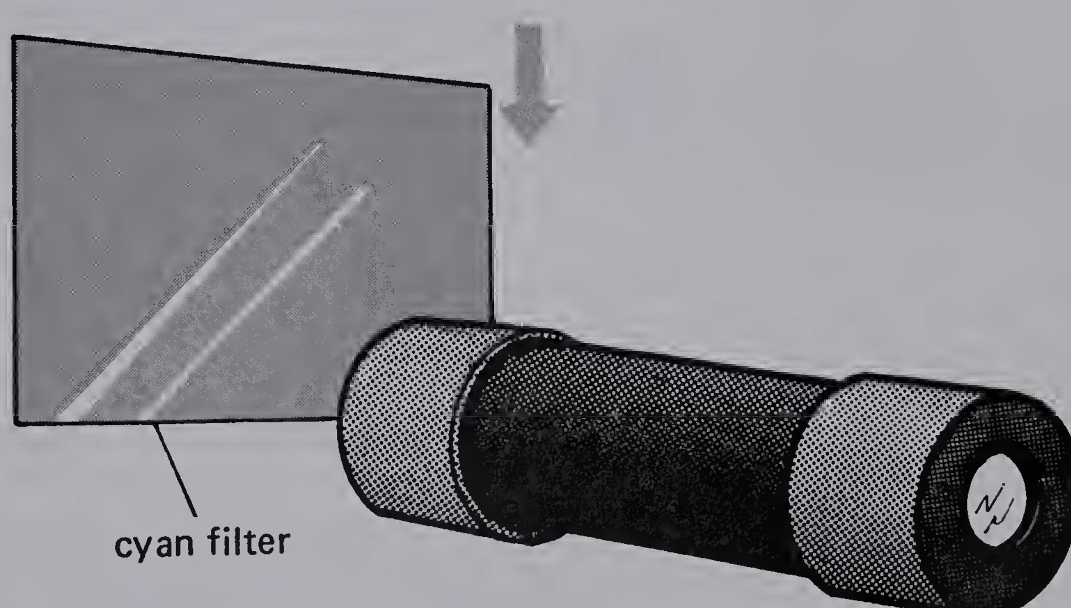


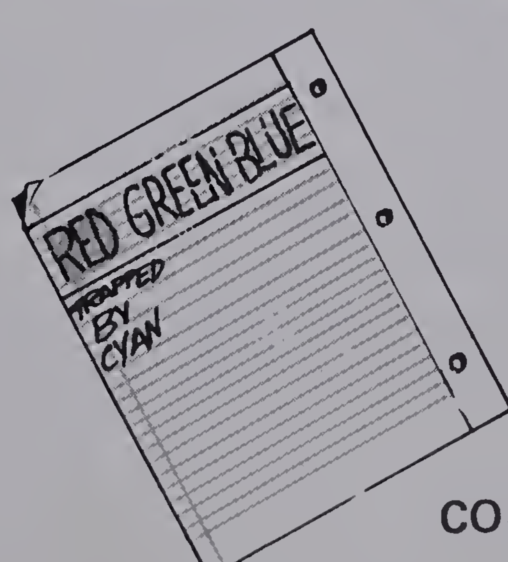
Figure 3-2

B. Get the cyan filter. Look through the spectroscope again. Slowly move the cyan filter in front of the spectroscope. Notice if there are any parts of the spectrum you can no longer see.



C. In your notebook, make a sketch showing which of the three colors — blue, green, or red — the cyan filter absorbed (trapped).

D. Now repeat Steps B and C, using the magenta filter and then the yellow filter. Be sure to record the colors of light each filter absorbed.



3-2. Yellow

3-3. Use the magenta and cyan filters together.

• 3-2. Which filter absorbs blue light energy?

• 3-3. How could you trap both red and green light energy?

Filters can trap some parts (colors) of visible light energy. And, in a similar way, some of the gases in the air trap certain kinds of energy from the sun. Figure 3-3 below shows this.

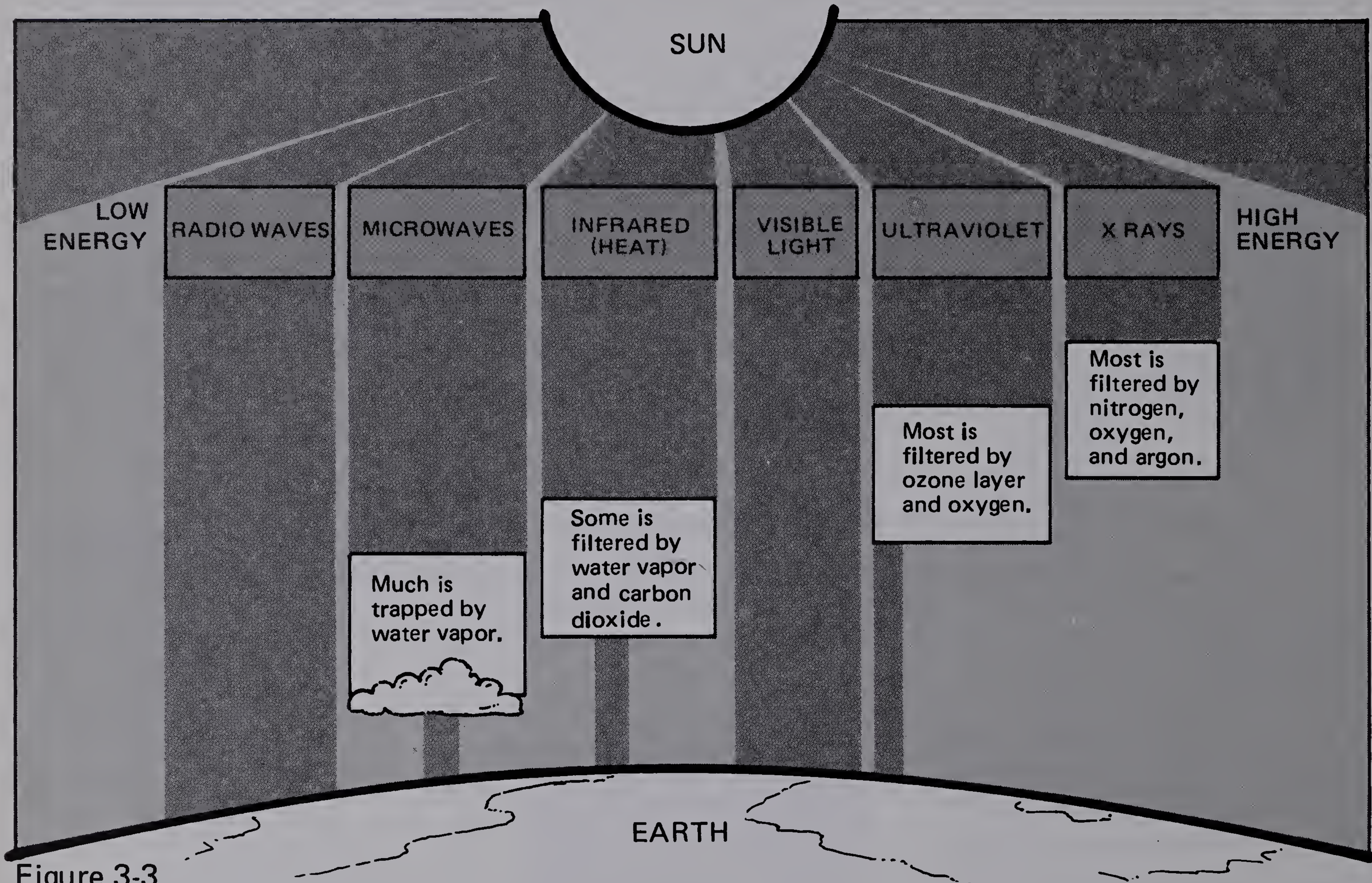


Figure 3-3

3-4. X rays and ultraviolet radiation; visible light and radio waves; infrared rays and microwaves

☆ **3-4. Which kinds of radiation are filtered out almost completely? Are not filtered? Are filtered somewhat?**

The sun's radiation ranges from very high-energy X rays to very low-energy radio waves. The high-energy X rays and ultraviolet waves are the kinds people worry about. If much of this high-energy radiation actually reached Earth, life as it is now could not exist. Fortunately, most of the x-radiation is trapped high above Earth by oxygen, nitrogen, and argon gases in the atmosphere.

The small amount of ultraviolet radiation that does get to Earth causes sunburn. If large amounts reached Earth, people could expect much more skin cancer as well as severe sunburn.

But ozone to the rescue! Ozone, O_3, is a form of oxygen. It is three oxygen atoms bonded together. Look at Figure 3-4 below.

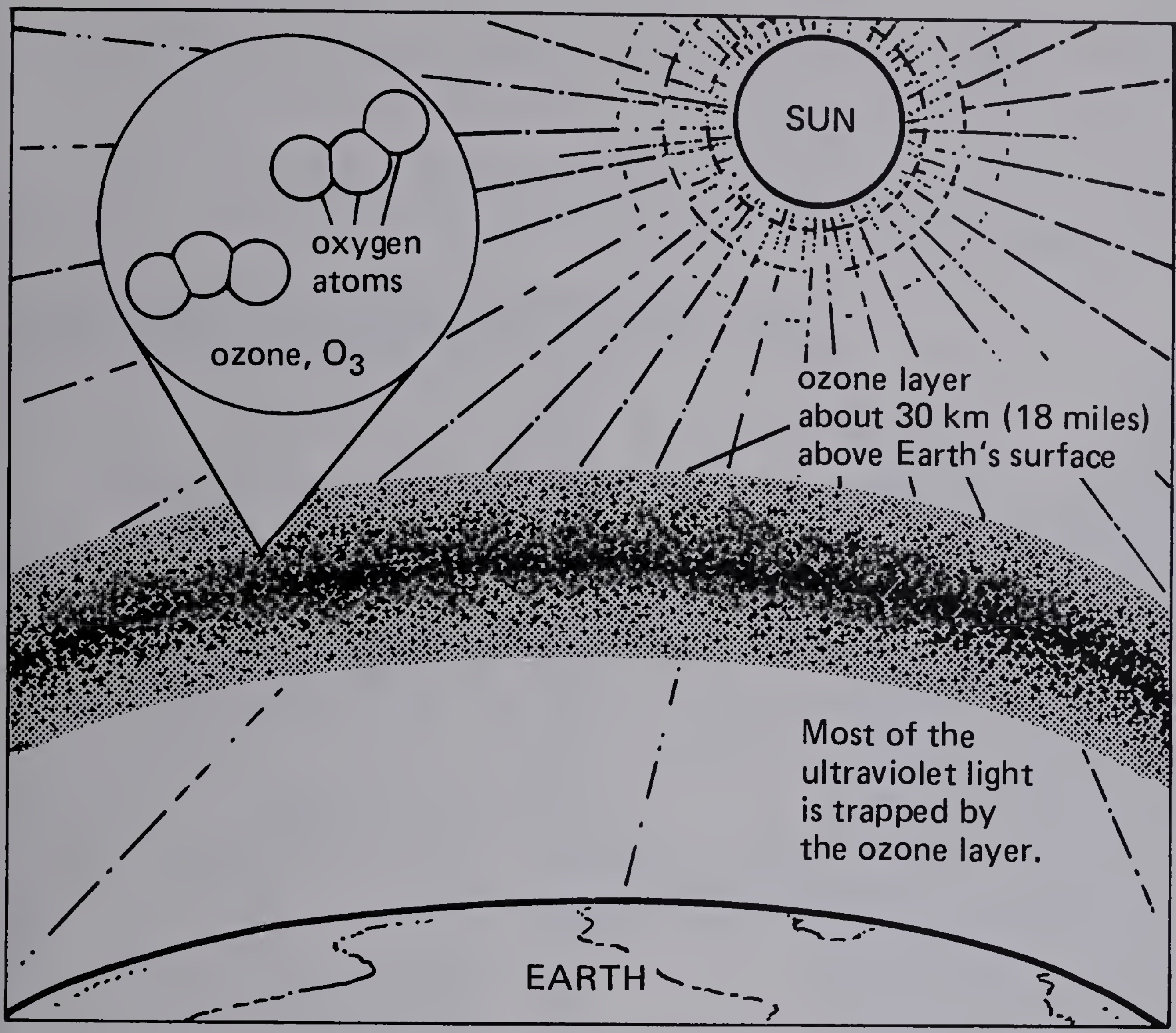


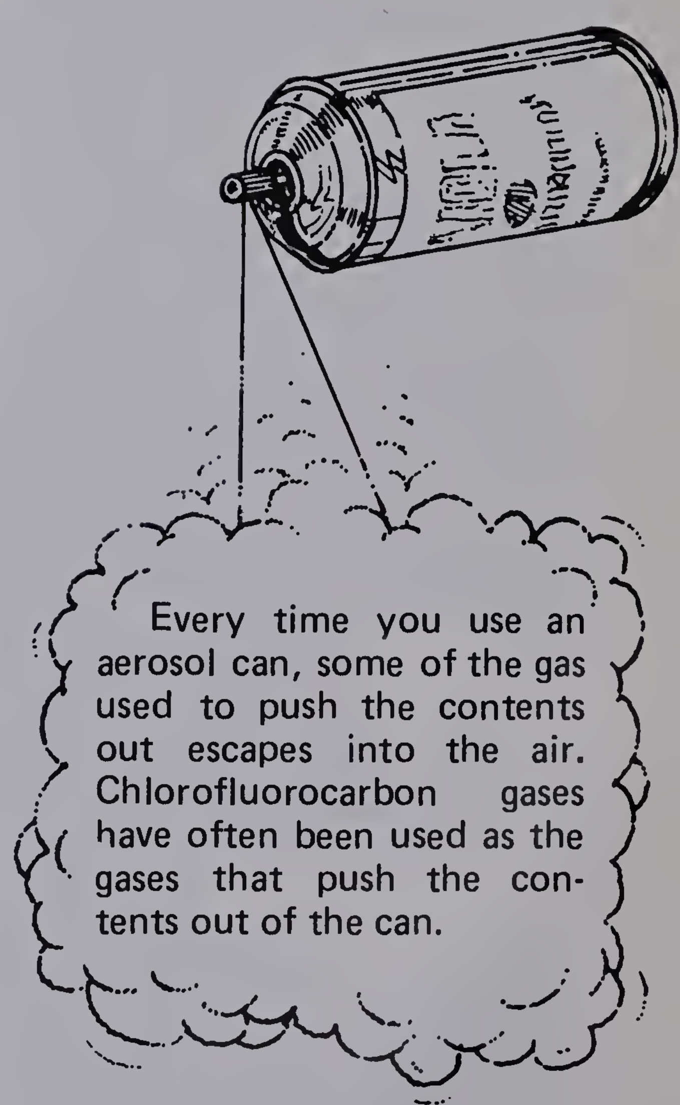
Figure 3-4

Chlorofluorocarbons are compounds of chlorine, fluorine, and carbon. They have been used in spray cans since 1944 and in refrigeration and air conditioning since 1931.

People never used to worry about letting these compounds into the air. The gases weren't known to react chemically with other substances. But scientists have now learned more about these gases. In the upper atmosphere, where ozone is, chlorofluorocarbons get enough energy from the sun to react. And there is evidence that these chemical reactions are destroying part of the ozone layer, letting more ultraviolet light reach Earth.

This was difficult to predict when such gases were first used in aerosol sprays. But now scientists are learning more about how chlorofluorocarbons react in the atmosphere. They are studying the problem very carefully.

As yet, nobody is sure just how serious the problem is. Scientists will have to learn more about chlorofluorocarbon—ozone reactions. Until they do, people should be cautious in using aerosol sprays that contain these gases.



3-5. Ultraviolet radiation

● 3-5. If chlorofluorocarbons are seriously disturbing the ozone layer, what kind of solar radiation reaching Earth may be increasing?

3-6. Scientists didn't have enough information on how these compounds reacted in the ozone layer.

☆ **3-6. Why was it difficult to predict that the use of chlorofluorocarbons in aerosol sprays might have harmful consequences?**

3-7. Severe sunburn and skin cancer

☆ **3-7. What could be the harmful effects on people if there was no ozone layer?**

ACTIVITY EMPHASIS: There are many sources of air pollution, but more than eighty percent of the human-caused pollution in the United States comes from the burning of fuels in transportation, stationary power generation, and industrial processes. The various pollution-detection methods depend on the particular properties of each pollutant.

MATERIALS PER STUDENT LAB GROUP: None

ACTIVITY 4: THE TELLTALE TRACE

Any substance that reduces the quality of the air you breathe is a pollutant. It's easiest to detect pollution when and where it happens. That's true for air pollution caused by people and for air pollution caused by nature. Figure 4-1 below lists the five major *natural* air pollutants and some of their sources.

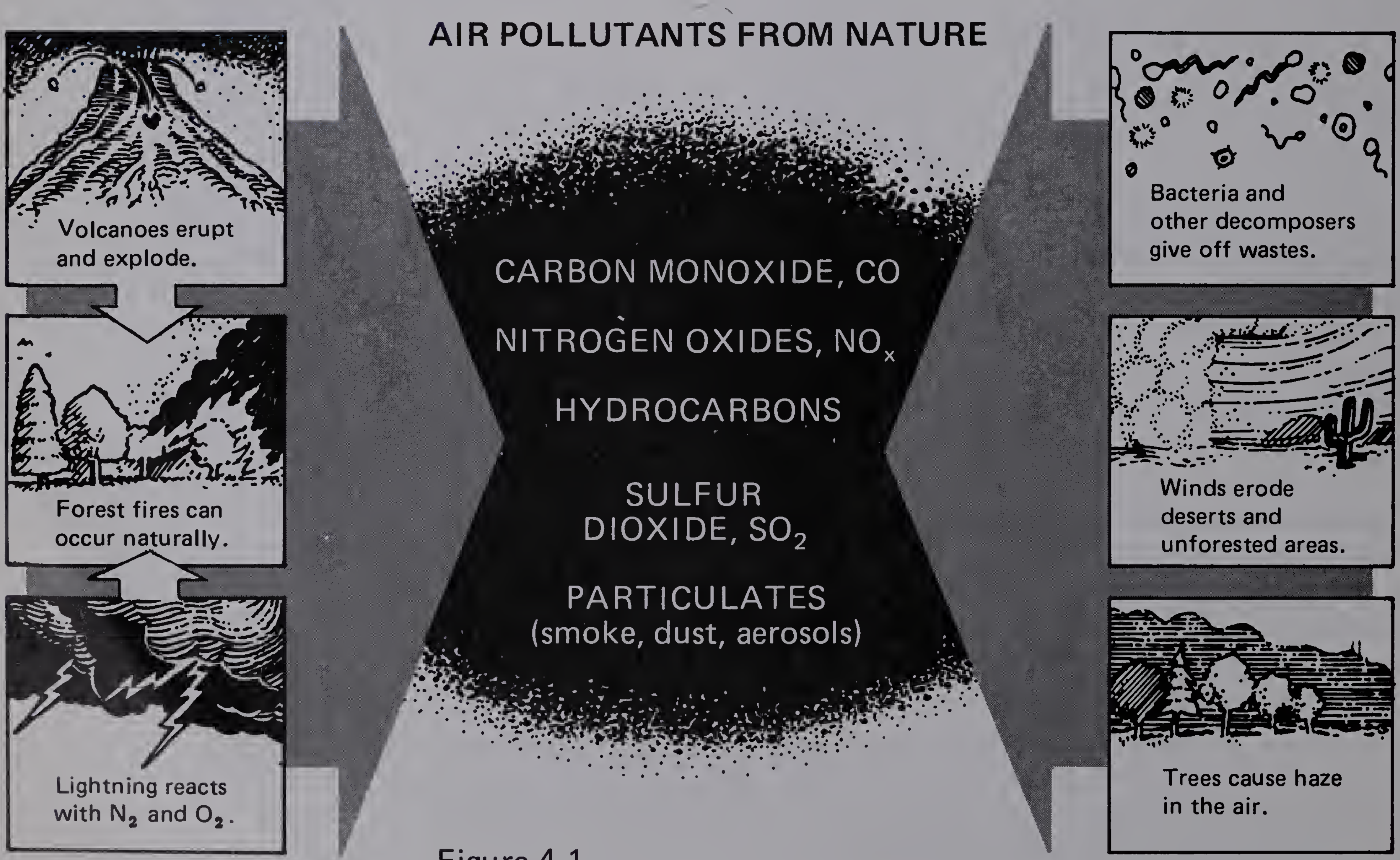


Figure 4-1

The main products of burning — water, H_2O, and carbon dioxide, CO_2 — are not classed as pollutants unless they're produced in huge amounts.

● 4-1. What is an air pollutant? What are the five major *natural* air pollutants?

4-1. An air pollutant is anything that reduces the quality of air. The major natural air pollutants are SO_2, particulates, CO, NO_x, and hydrocarbons.

Look at Figure 4-2 below. It lists the five major air pollutants caused by people and some sources of those pollutants.

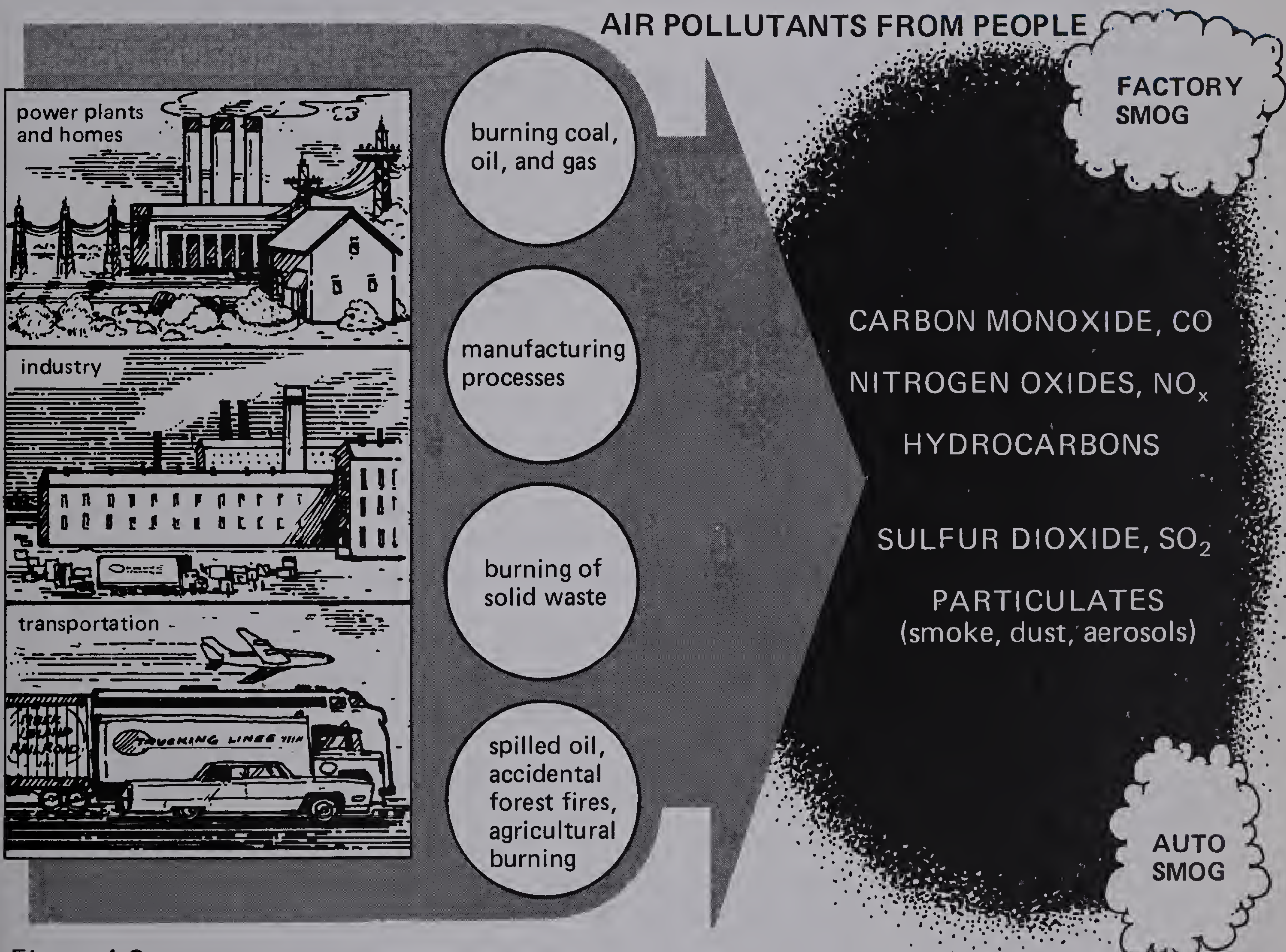


Figure 4-2

Figure 4-1 (page 14) represents from ten to a hundred times more pollution than Figure 4-2 above. But you've probably heard more about the pollution caused by people.

● 4-2. Are the major pollutants caused by people the same ones as the major natural pollutants? Why do you think you hear more about those caused by people?

4-2. Yes; they are put into the air near where most people live or work by factories, cars, and homes. Most cities have smog that's easy to see. These pollutants bother people more, since they're around them.

Smog is a sure sign that people have polluted.

In old cities with narrow streets that have few cars but many industries, power plants, and apartment houses, coal burning produces coal smog. It irritates the throat and lungs.

In new cities with many cars and single-family homes, the burning of gasoline, natural gas, and oil produces auto smog. It is also irritating, especially to the eyes.

In an industrial city surrounded by new suburbs and crossed by highways, there might be both kinds of smog all year. That's why the government issues air-pollution warnings. And it may also enforce manufacturing slowdowns.

4-3. By the irritation of eyes, throat, and lungs

• 4-3. What is one way that people can detect air pollution?

Both kinds of smog irritate the eyes or the throat and lungs. That's one way to detect pollutants. But it's not very accurate. There are several very accurate ways to detect both the kinds and amounts of pollutants in air. And these have to be used all the time to avoid air pollution that could seriously harm people.

All these pollution-detection methods depend on the particular properties of the different pollutants. For example, solids, liquids, and gases are detected by different methods. Particulates are detected in one way, gaseous pollutants in other ways.

4-4. They all depend on the differences in properties among different air pollutants.

☆ **4-4. What do all the different methods of detecting pollutants depend on?**

Look at Figure 4-3 below. It shows two methods of detecting and measuring particulates — soot, dust, and so on.

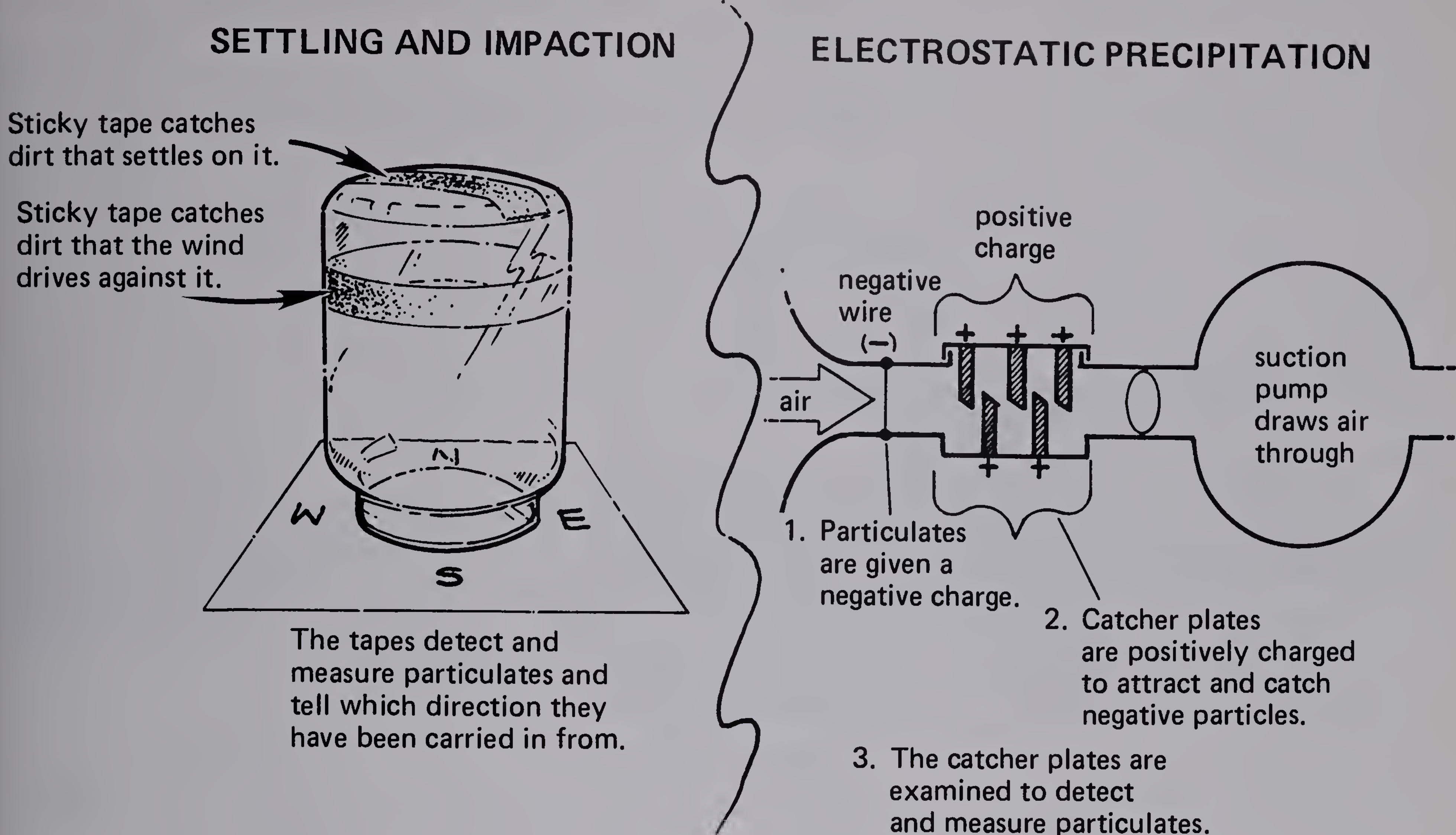


Figure 4-3

• 4-5. Why do you think the methods in Figure 4-3 above work for particulates but not for gases?

4-5. Gases don't stick to tape (or the tape would be nonsticky right away). Dirt does stick. Charged particles stick to the catcher plate, but gases don't.

Now for the gaseous pollutants. Carbon compounds have special properties that let them be detected. Certain metals speed up the reaction of gaseous carbon compounds with oxygen. The heat of this fast reaction can make a thin wire of the metal glow brightly. Look at Figure 4-4 below.

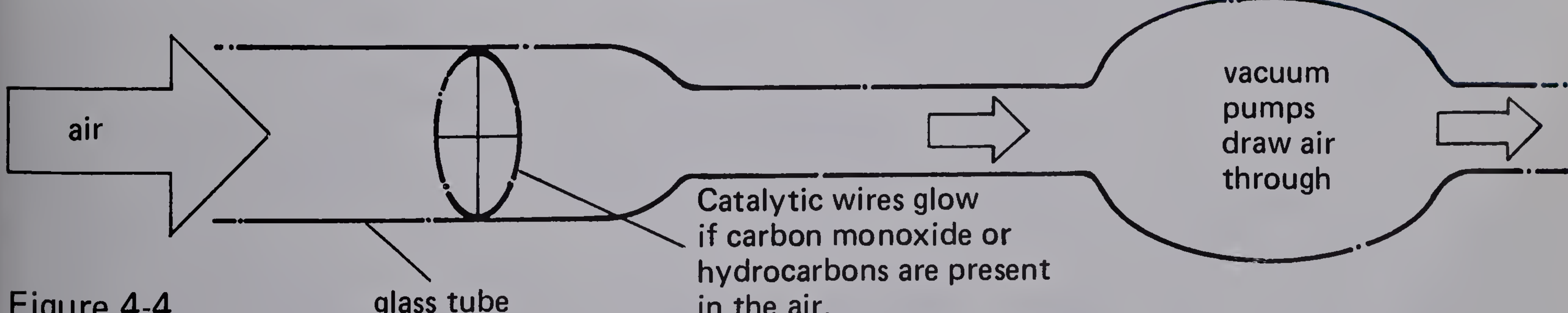


Figure 4-4

This method can be used to tell whether carbon monoxide or hydrocarbons are present in air, chimney smoke, or the exhaust from a car or airplane engine.

One way to detect every kind of gaseous air pollutant depends on the fact that different gases absorb light differently. Figure 4-5 below shows how pollutant gases are detected by their different patterns of absorbing light.

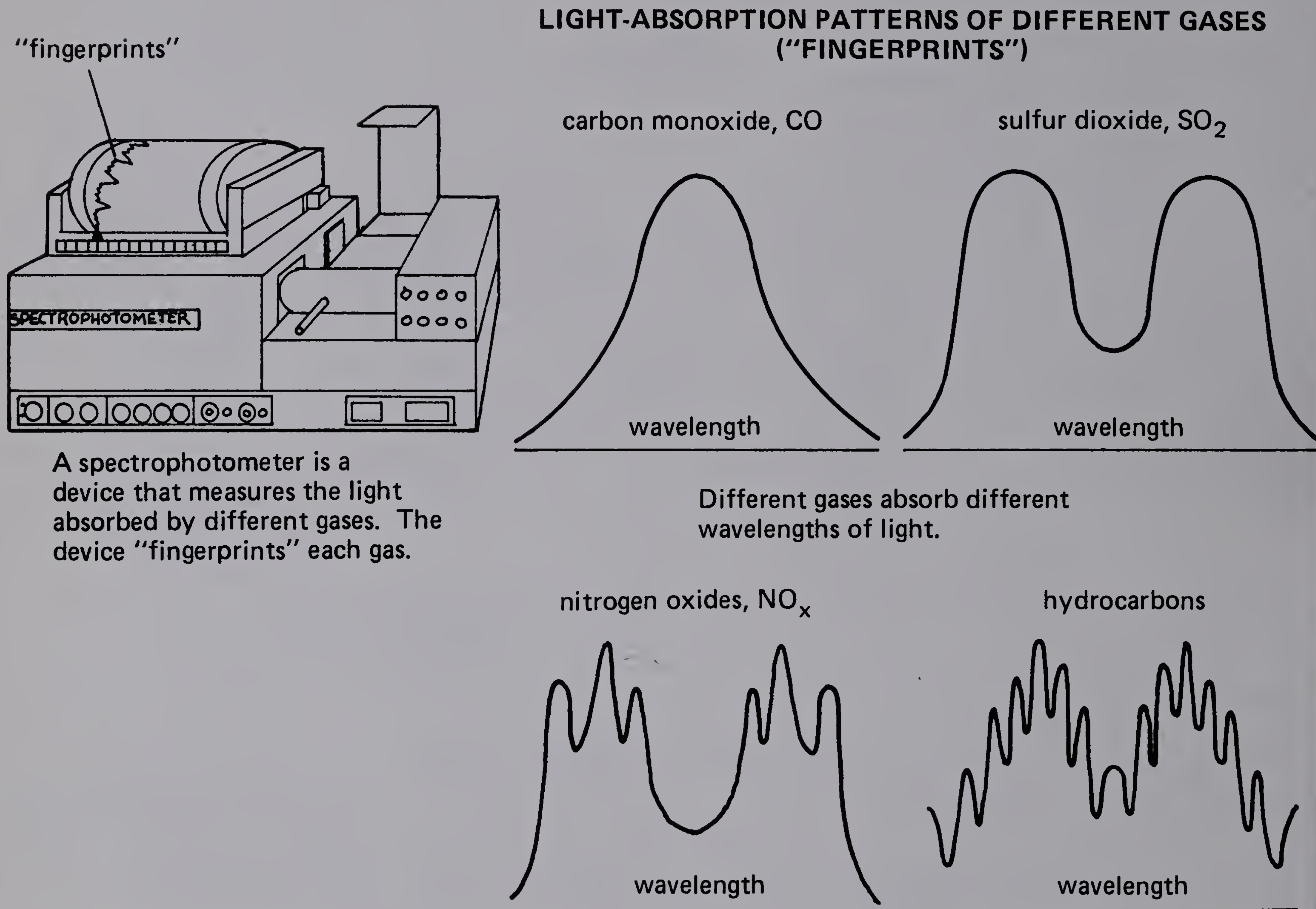


Figure 4-5

4-6. They depend on differences in the absorption (filtering) of light by different gases.

• 4-6. What do the "fingerprint" methods of gas detection depend on?

If there are enough gaseous pollutants in the air, they can be detected immediately by this method. If not, they have to be separated from air and sometimes from each other for "fingerprinting."

4-7. Settling and impaction (collect them on sticky tape) or electrostatic precipitation (give them a negative charge and collect them on positively charged catcher plates)

☆ **4-7. Describe two ways to accurately detect air pollutants that are particulates.**

4-8. Measure light-absorption patterns; every gas absorbs light differently.

☆ **4-8. Describe a method for detecting any gaseous air pollutant. What property does this method depend on?**

ACTIVITY 5: HELP FOR NATURE

ACTIVITY EMPHASIS: Nature's cleanup of the five major pollutants can't take care of increased human pollution. The increase can be controlled or prevented using exact knowledge of each kind of pollution.

Nature can remove from the air almost every known pollutant. Natural pollution control kept the air clean for a long time. Figure 5-1 below shows some of the ways nature cleans the air.

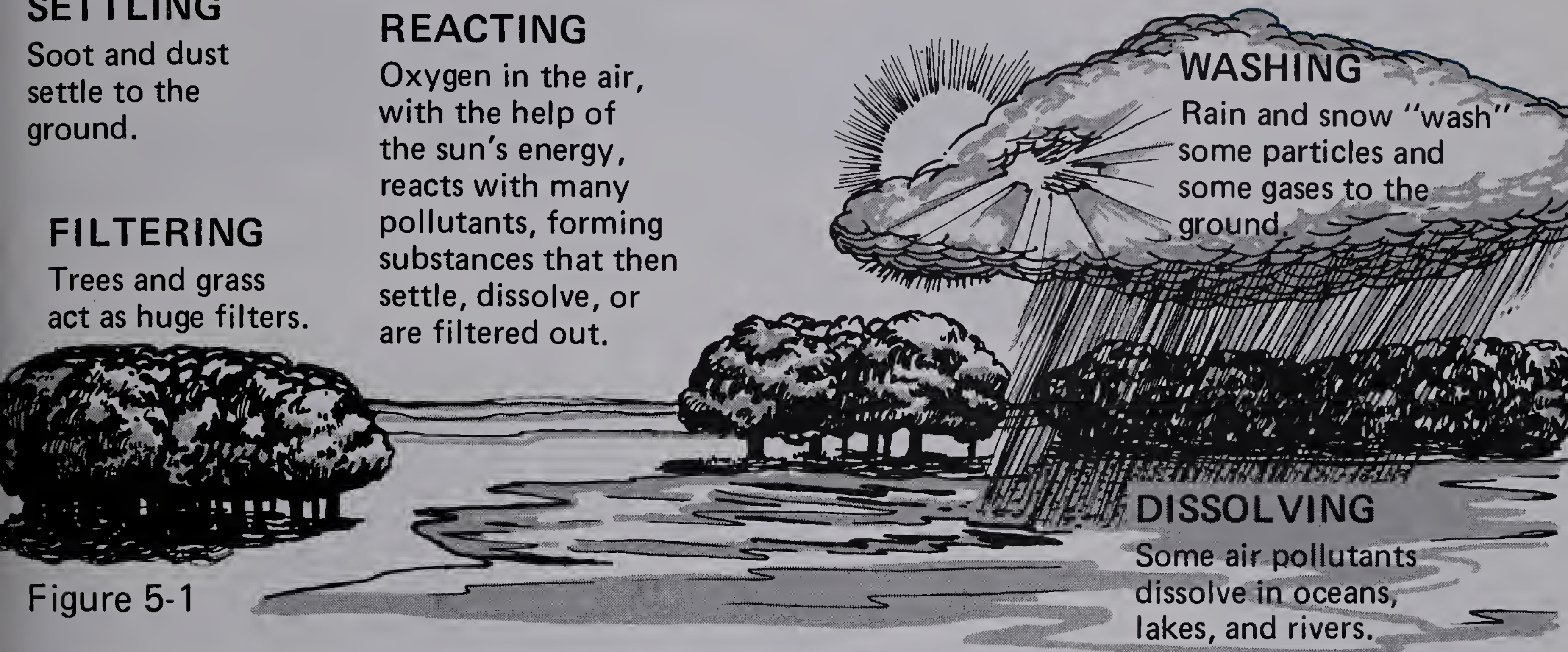


Figure 5-1

MATERIALS PER STUDENT LAB GROUP: None

But nowadays people have to help nature. Industries, cars, and power plants are dumping pollutants into the air much faster than natural processes can handle them.

Reacting now produces smogs, which burn the eyes and irritate the throat and lungs. *Washing* and *dissolving* now produce acid rains, which pit building stones and kill plants. *Settling* now blackens cities and the lungs of people who live in them.

Burning produces a very large percentage of the total air pollution. Fuels are burned in cars, in power plants, and in factories. Figure 5-2 below shows the five major pollutants that are produced by burning.

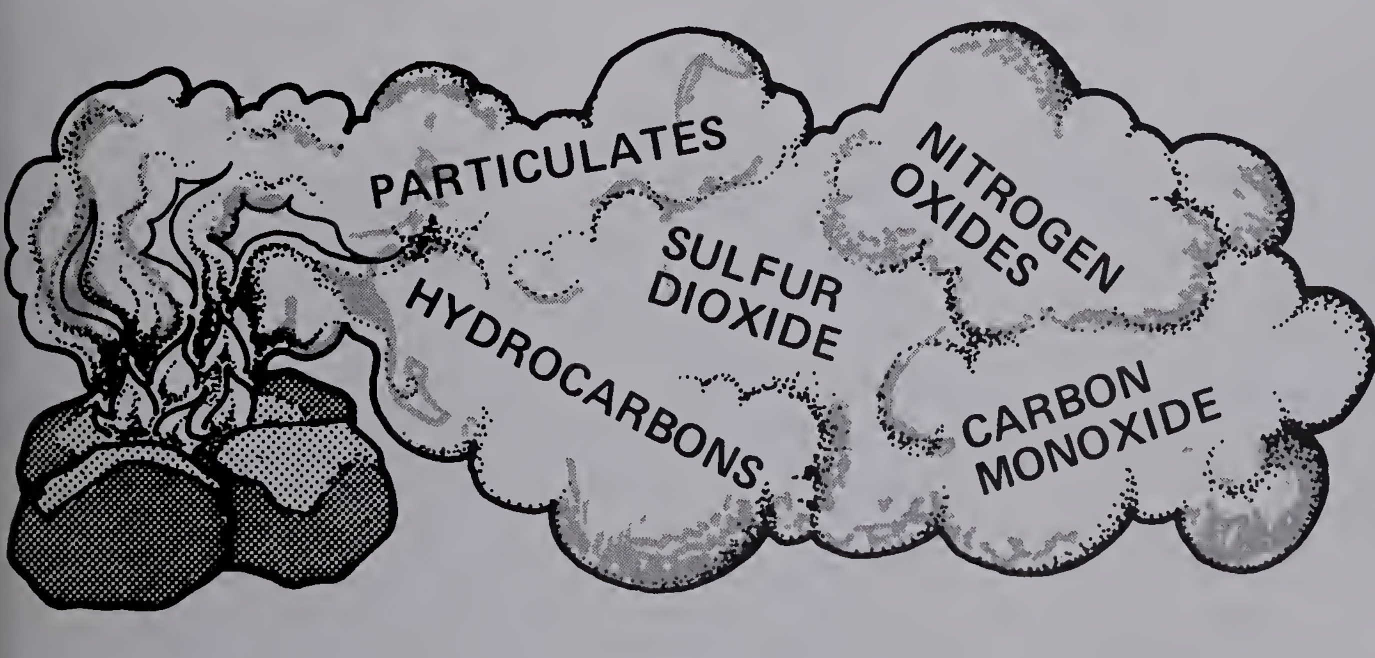


Figure 5-2

Take a close look at how each of the five major pollutants is formed. Then you can get some clues about how to control each pollutant. Look first at particulates — one of the easiest to control.

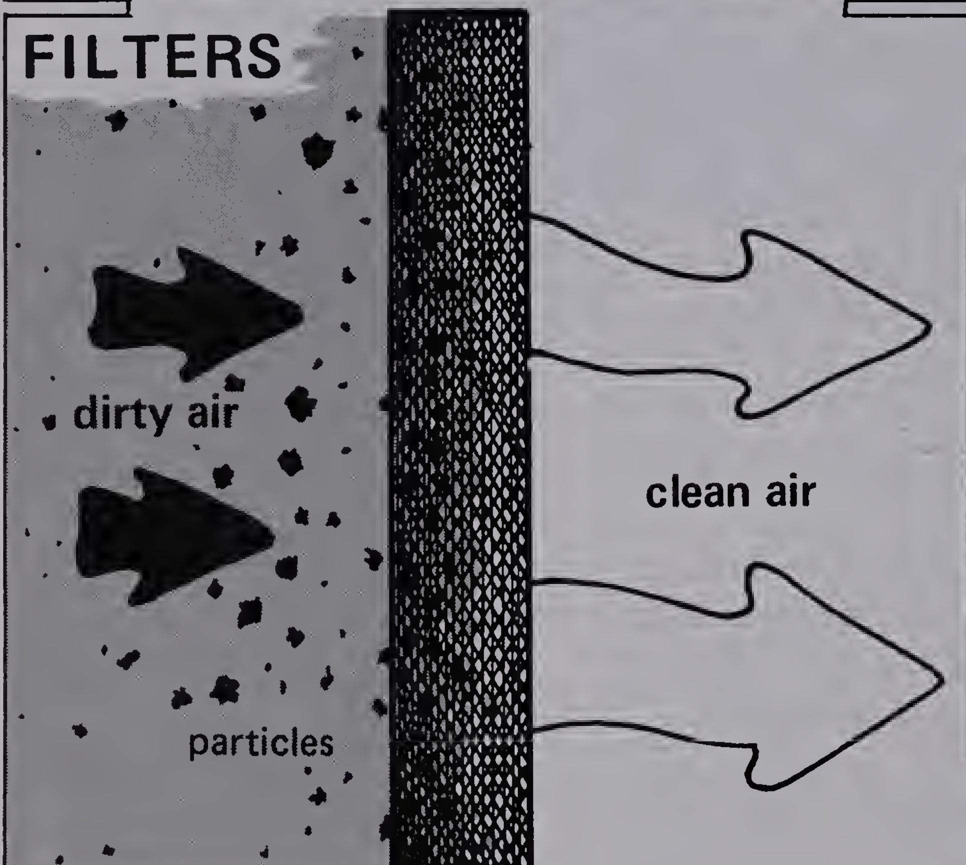
PARTICULATES

When coal, wood, and other solids burn, **powdery materials**, such as **soot** and **dust**, are carried into the air. **Tiny particles** also come from mining operations, factories, and spraying.

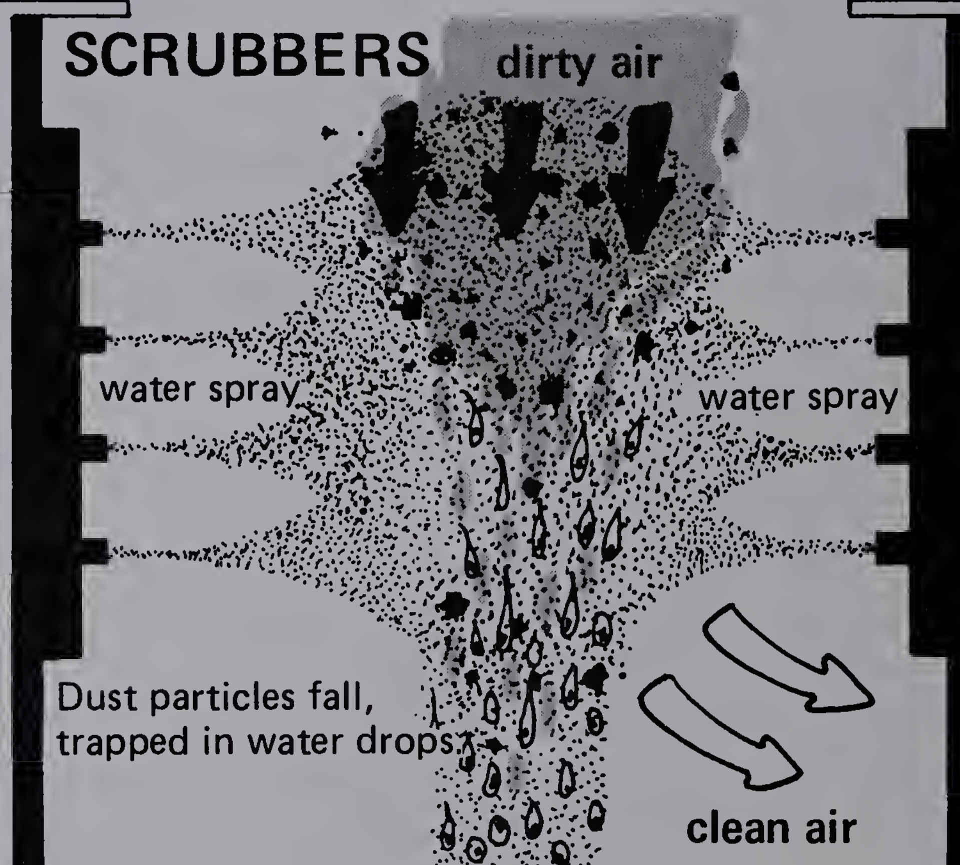
CONTROL

Factories and power plants can install devices that trap particulates.

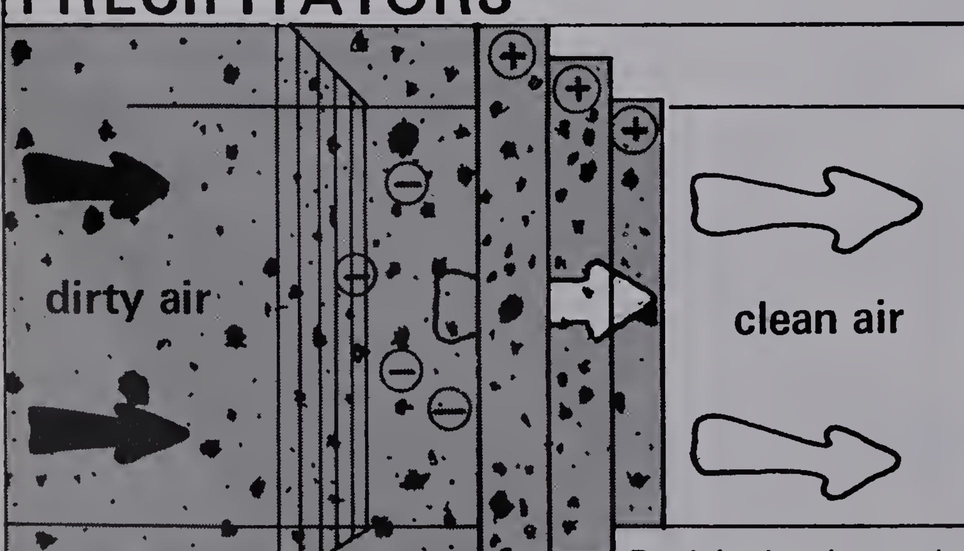
FILTERS



SCRUBBERS



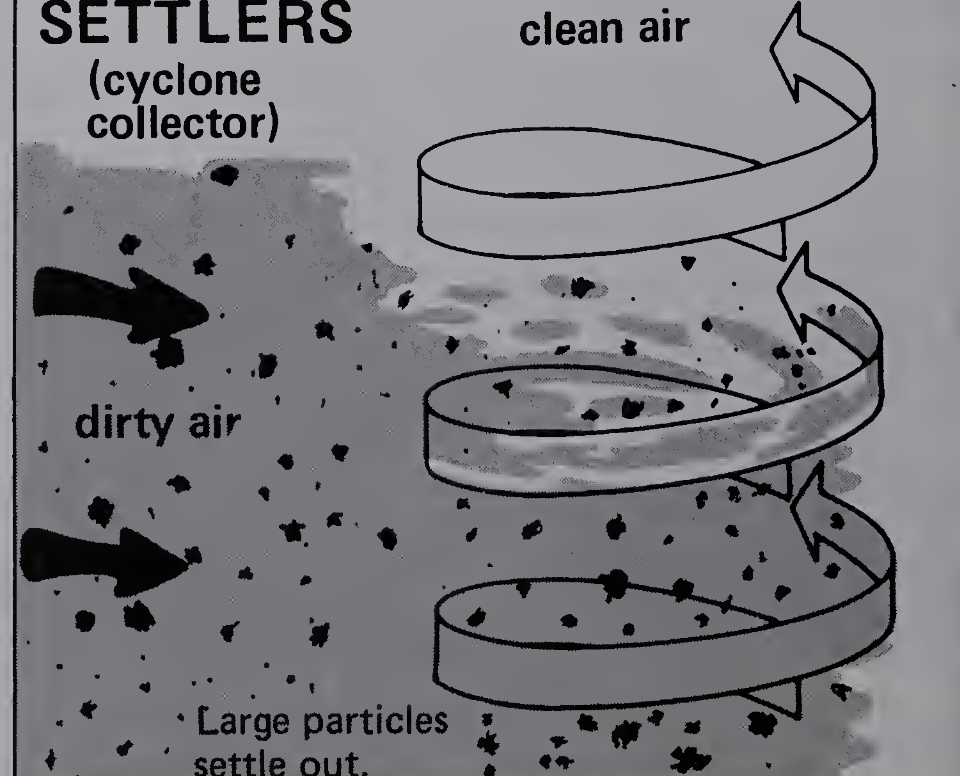
ELECTROSTATIC PRECIPITATORS



Negatively charged wires give dirt particles a negative charge.

Positively charged plates attract and collect negatively charged particles.

SETTLERS (cyclone collector)



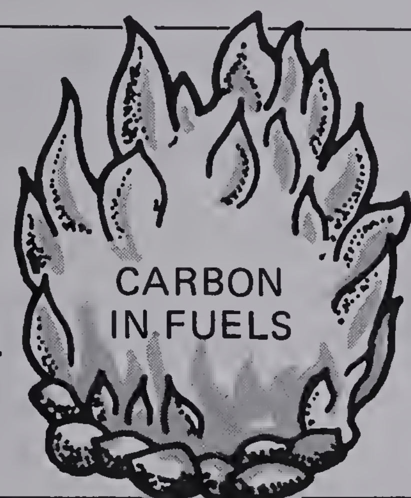
• 5-1. Suppose an electrical power plant burns coal, producing large amounts of gray-black smoke. What devices might the plant use to reduce particulates from its smokestacks?

5-1. Devices for filtering, scrubbing, precipitating, or settling out particulates

The other four major pollutants — carbon monoxide, hydrocarbons, nitrogen oxides, and sulfur dioxide — are gases. The first three come mostly from the use of cars. They're harder to control than particulates. Again, clues for controlling them come from the way they're formed.

CARBON MONOXIDE, CO

If there's not enough air for a fuel to burn completely, carbon monoxide, CO, is formed. It's a colorless gas that's a deadly poison.



PLENTY OF AIR (oxygen) $\xrightarrow[\text{complete burning}]{\text{higher temperatures}}$ CO_2 (carbon dioxide) NOT POISONOUS

NOT ENOUGH AIR (oxygen) $\xrightarrow[\text{incomplete burning}]{\text{lower temperatures}}$ CO (carbon monoxide) POISONOUS

CONTROL

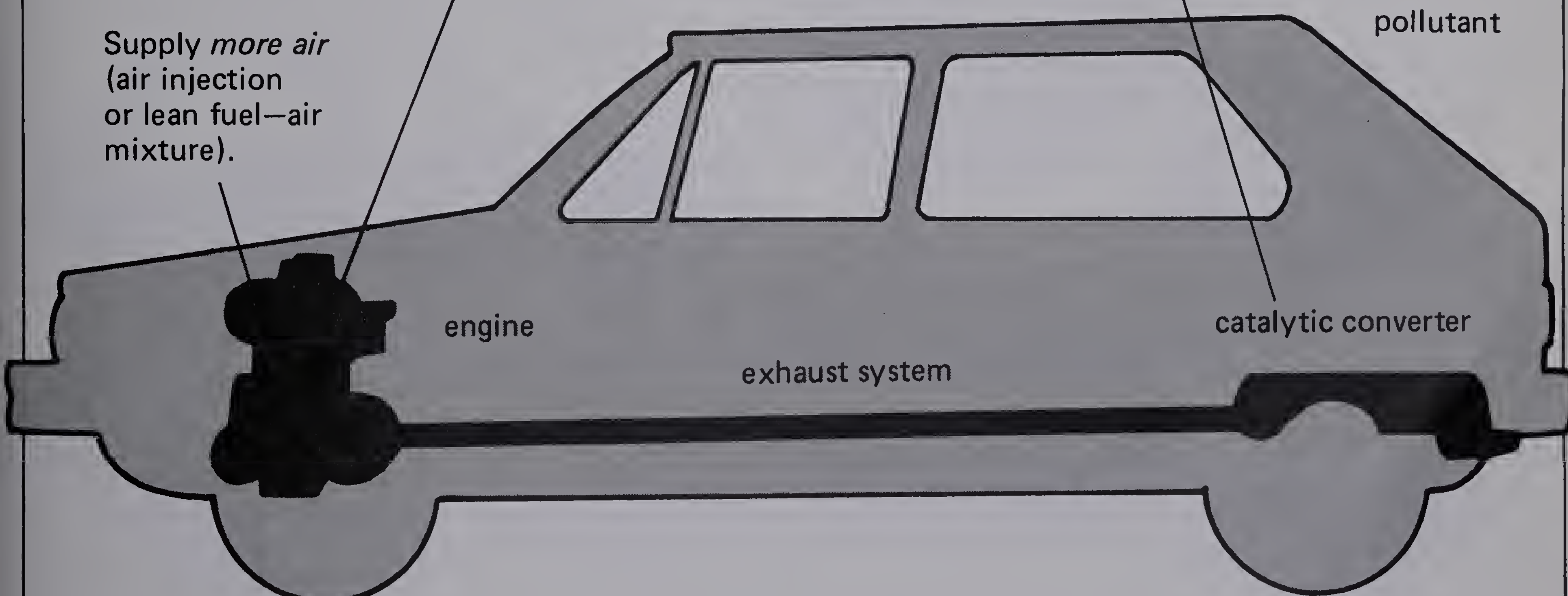
Burn the fuel completely.

If CO still forms, burn it in a catalytic converter. Use *high temperatures.*

$$\underset{\text{pollutant}}{2\,CO + O_2} \xrightarrow[\text{more}]{\text{burn}} \underset{\text{not a pollutant}}{2\,CO_2}$$

For incompletely burned fuels, mix them with *more air* and recycle them through burning chamber.

Supply *more air* (air injection or lean fuel—air mixture).



5-2. CO forms when there's not enough air or high enough temperatures, causing incomplete burning.

5-3. Supply more air (with air injection or a leaner fuel—air mixture) and recycle incompletely burned fuel back into the burning chamber. Burn CO to CO_2 at high temperatures in a catalytic converter.

- 5-2. What causes carbon monoxide, CO, to form during burning?

- 5-3. To reduce CO pollution from a car, what can be done in the engine? In the exhaust system?

Hydrocarbons are compounds of hydrogen and carbons. Gasoline, fuel oil, and natural gas are mostly hydrocarbon compounds.

HYDROCARBONS

When fuels evaporate, pollutants called *hydrocarbons* are put into the air.

Hydrocarbons also result when fuels are not burned completely.

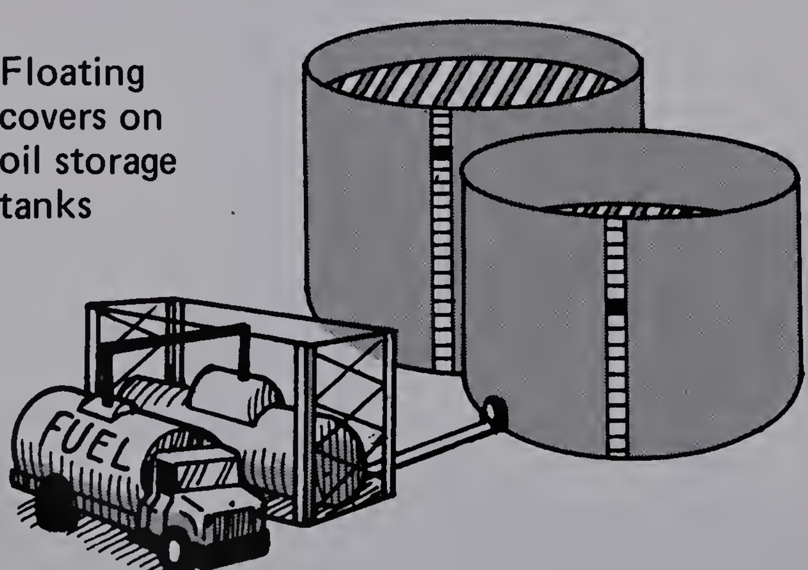
PLENTY OF AIR (oxygen) $\xrightarrow[\text{complete burning}]{\text{higher temperatures}}$ $CO_2 + H_2O$ (carbon dioxide) (water)

NOT ENOUGH AIR (oxygen) $\xrightarrow[\text{incomplete burning}]{\text{lower temperatures}}$ hydrocarbons + CO (unburned) (carbon monoxide)

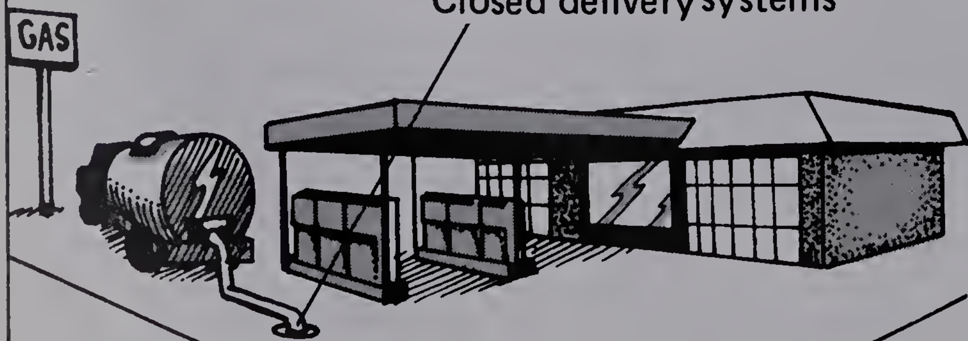
CONTROL

CONTROL OF EVAPORATION

Floating covers on oil storage tanks



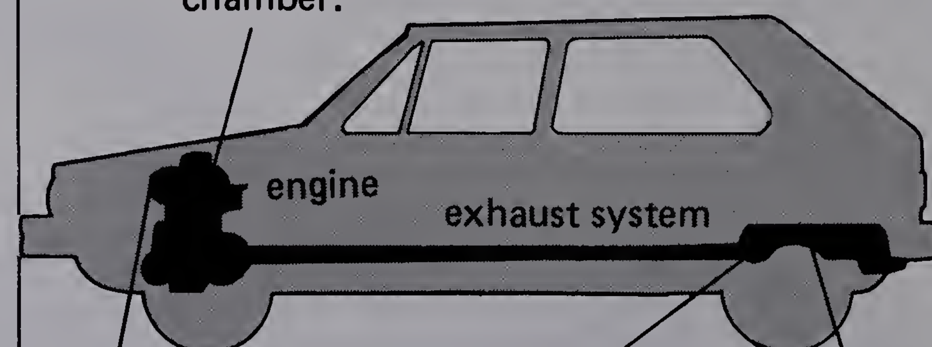
Closed delivery systems



CONTROL OF INCOMPLETE BURNING

Unburned hydrocarbons are produced by conditions that produce CO and controlled by methods that control CO.

Recycle unburned hydrocarbons back through the burning chamber.



hydrocarbons + $O_2 \rightarrow CO_2 + H_2O$

Supply *more air* to burn them completely.

Catalytic converters burn the rest of the hydrocarbons at *high temperatures.*

• 5-4. How can hydrocarbon pollution from the evaporation of liquid fuels be reduced?

5-4. Use floating tank covers and closed delivery systems.

• 5-5. Describe three ways to reduce hydrocarbon pollution from cars.

5-5. Supply sufficient air for complete burning. Recycle incompletely burned hydrocarbons through the burning chamber. Burn unburned hydrocarbons in catalytic converters in the exhaust system.

• 5-6. What two pollutants from cars require the same kinds of control procedures?

5-6. Carbon monoxide and hydrocarbons

Luckily, the same control procedures can be used on cars for both hydrocarbons and carbon monoxide. But, unfortunately, these same procedures tend to increase the amount of the fourth kind of pollutant — nitrogen oxides.

NITROGEN OXIDES, NO_x (number varies)

When fuels are burned at high temperatures with lots of air, two gases from the air around the burning fuel combine.

$$\underset{\text{(nitrogen)}}{N_2} + \underset{\text{(oxygen)}}{O_2} \xrightarrow[\text{temperature}]{\text{high}} \underset{\substack{\text{(nitrogen oxides)}\\ \text{pollutants}}}{NO_x}$$

CONTROL

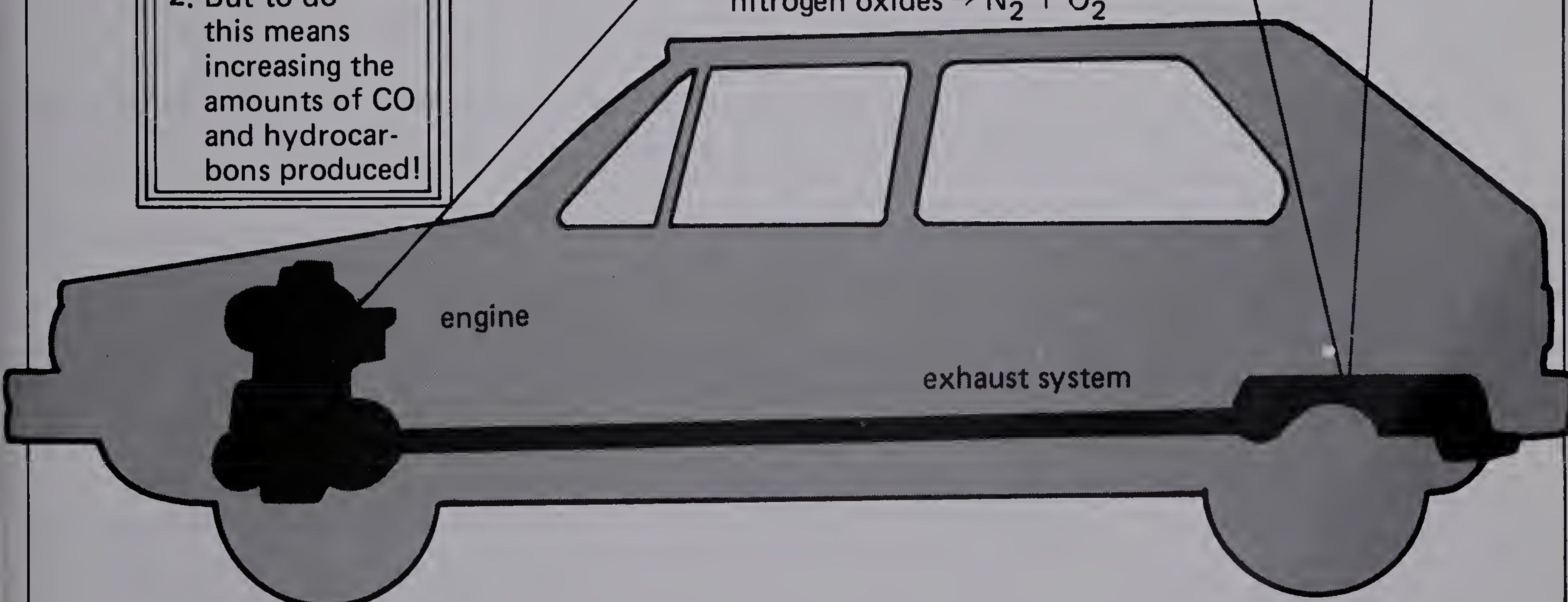
1. When burning fuels, lower the temperatures and reduce the amount of air.

2. But to do this means increasing the amounts of CO and hydrocarbons produced!

3. The solution is to keep devices for decreasing CO and hydrocarbons.

Then add a second catalytic converter to break the nitrogen oxides into harmless nitrogen and oxygen gases again.

nitrogen oxides $\rightarrow N_2 + O_2$



-7. They are best controlled by using a second catalytic converter to convert NO_x into N_2 and O_2. Lowering the temperature and the air supply would increase CO and hydrocarbons (and even particulates).

• 5-7. How are polluting nitrogen oxides best controlled in cars? Why can't the temperature and air supply in the engine just be lowered?

The fifth major air pollutant is sulfur dioxide. Though most fuels are mainly carbon and hydrogen, many of them have sulfur impurities in them. When these fuels are burned, sulfur dioxide gas is formed. This gas is poisonous to plants and animals.

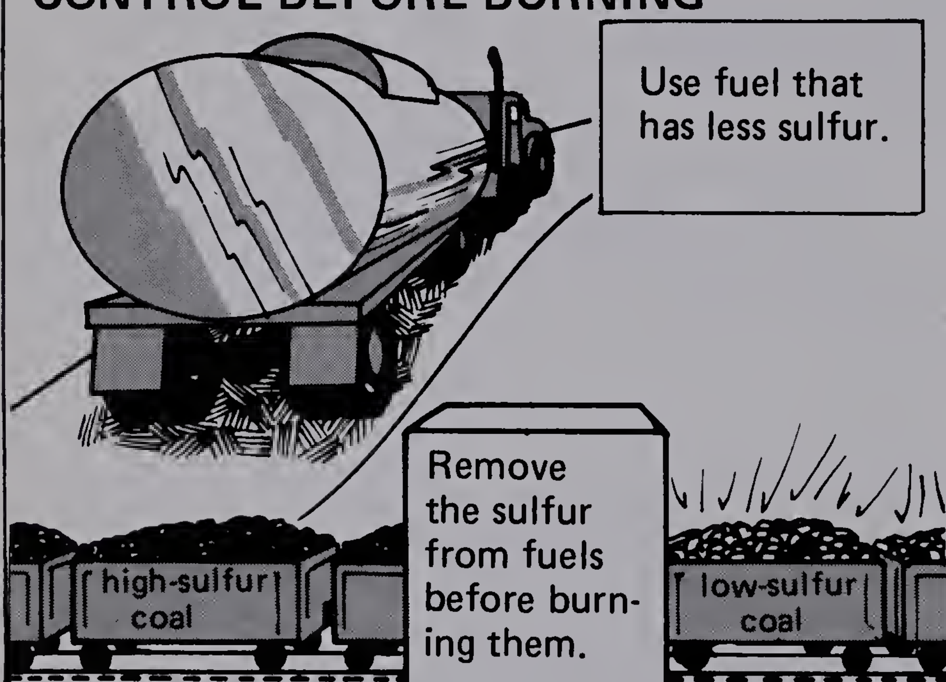
Coal has the most sulfur, but oil, gasoline, and natural gas have some sulfur too. The production of sulfur dioxide from these fuels, especially coal, is controlled both before and after burning.

SULFUR DIOXIDE, SO_2

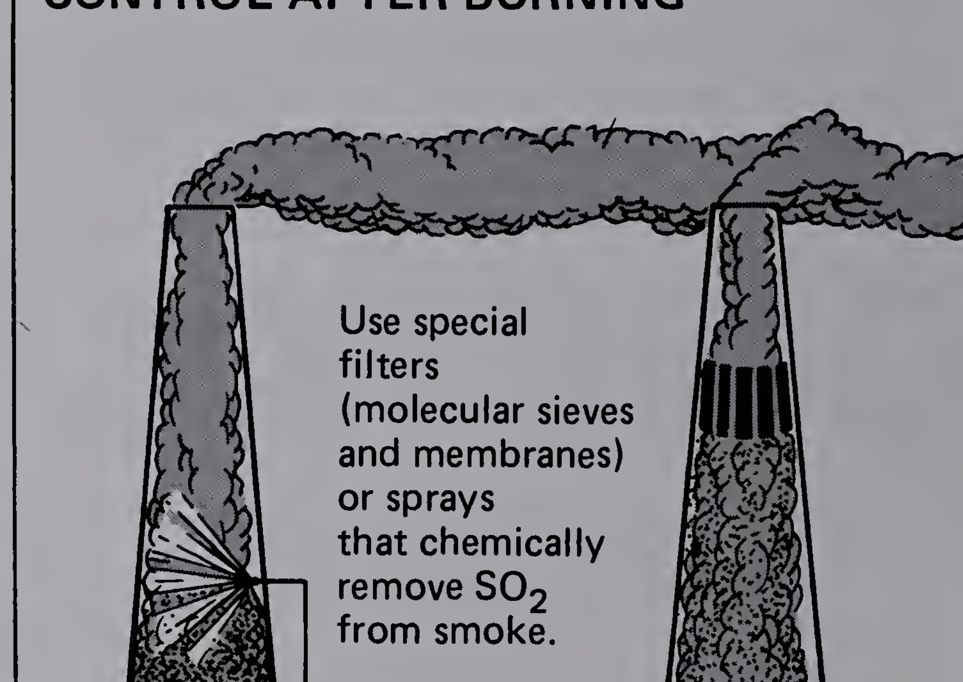
In the burning of fuels, sulfur forms sulfur dioxide.

$$S + O_2 \rightarrow SO_2$$

CONTROL BEFORE BURNING



CONTROL AFTER BURNING



5-8. By chemical removal and by special filters

• 5-8. In what two ways can sulfur dioxide be removed from smokestack gases after fuel has been burned?

5-9. Remove sulfur from fuels or switch to fuels that have less sulfur.

• 5-9. What can be done before burning that will lower the amount of SO_2 entering the air from burning fuels?

5-10. A4, 6; B1, 2; C1, 2, 5; D2; E3, 6

☆ **5-10. Match each pollutant with one or more methods that can be used to control it.**

Pollutant	Method
A. Particulates	1. burn at higher temperatures with more air
B. Carbon monoxide, CO	2. use a catalytic converter
C. Hydrocarbons	3. switch to a low-sulfur fuel
D. Nitrogen oxides, NO_x	4. use an electrostatic precipitator
E. Sulfur dioxide, SO_2	5. prevent evaporation of fuel
	6. use filters in the smokestack

ACTIVITY 6: NAME YOUR POISON

Air pollution hurts all living things. Different pollutants and combinations of them have different effects on animals and plants.

The five major pollutants are carbon monoxide, particulates, nitrogen oxides, sulfur dioxide, and hydrocarbons. Each has harmful effects. And some of them, when combined to form smogs, are even more dangerous. (Smog is *sm*oke plus f*og*.)

Pollutants affect animals, including human beings, in two general ways. Some affect the ability of the blood to carry oxygen. Others irritate and damage parts of the respiratory system.

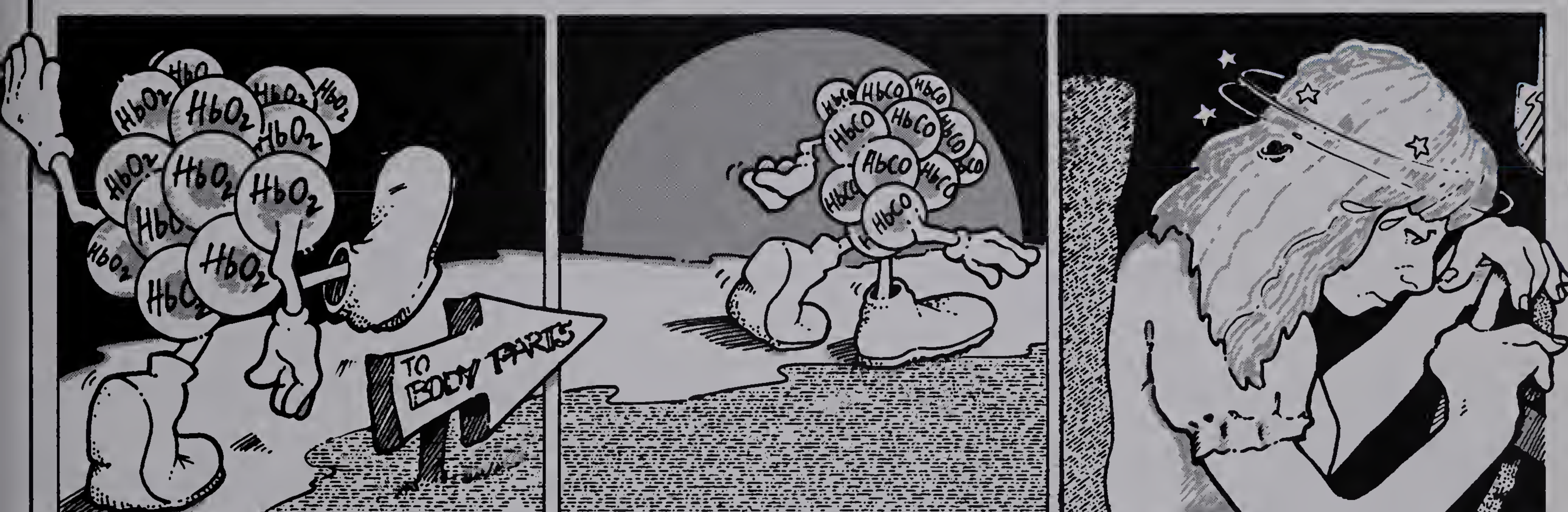
Carbon monoxide is deadly to animals. It has no odor, color, or taste, and it can kill you before you know it. It kills by keeping the body from getting oxygen. Look at Figure 6-1 below.

ACTIVITY EMPHASIS: By themselves and as changed in smog, the principal air pollutants are harmful to humans and other animals. They damage the respiratory system or reduce the oxygen level in the blood. Plants are also damaged, but they differ from animals in their sensitivity to various pollutants.

MATERIALS PER STUDENT LAB GROUP: None

The combined effect in smog is called a *synergistic effect*.

CARBON MONOXIDE POISONING



Hemoglobin, Hb, in red blood cells normally carries oxygen, O_2, to all body parts.	But hemoglobin combines with carbon monoxide, CO, about 200 times more readily than with oxygen, O_2. So, if CO is in air taken into the lungs, hemoglobin combines with CO instead of O_2.	CO keeps the hemoglobin from performing its vital function of carrying O_2 to body cells. Headaches, dizziness, nausea, and even death can result.

Figure 6-1

● 6-1. Any animal that doesn't get enough oxygen dies of suffocation. How does carbon monoxide keep the body from getting oxygen?

6-1. Hemoglobin combines with carbon monoxide. That reduces the ability of the Hb to carry oxygen to body cells.

Carbon monoxide reduces the oxygen-carrying capacity of the blood. The other major air pollutants irritate and damage the tissues of the respiratory system. Look at Figure 6-2 below.

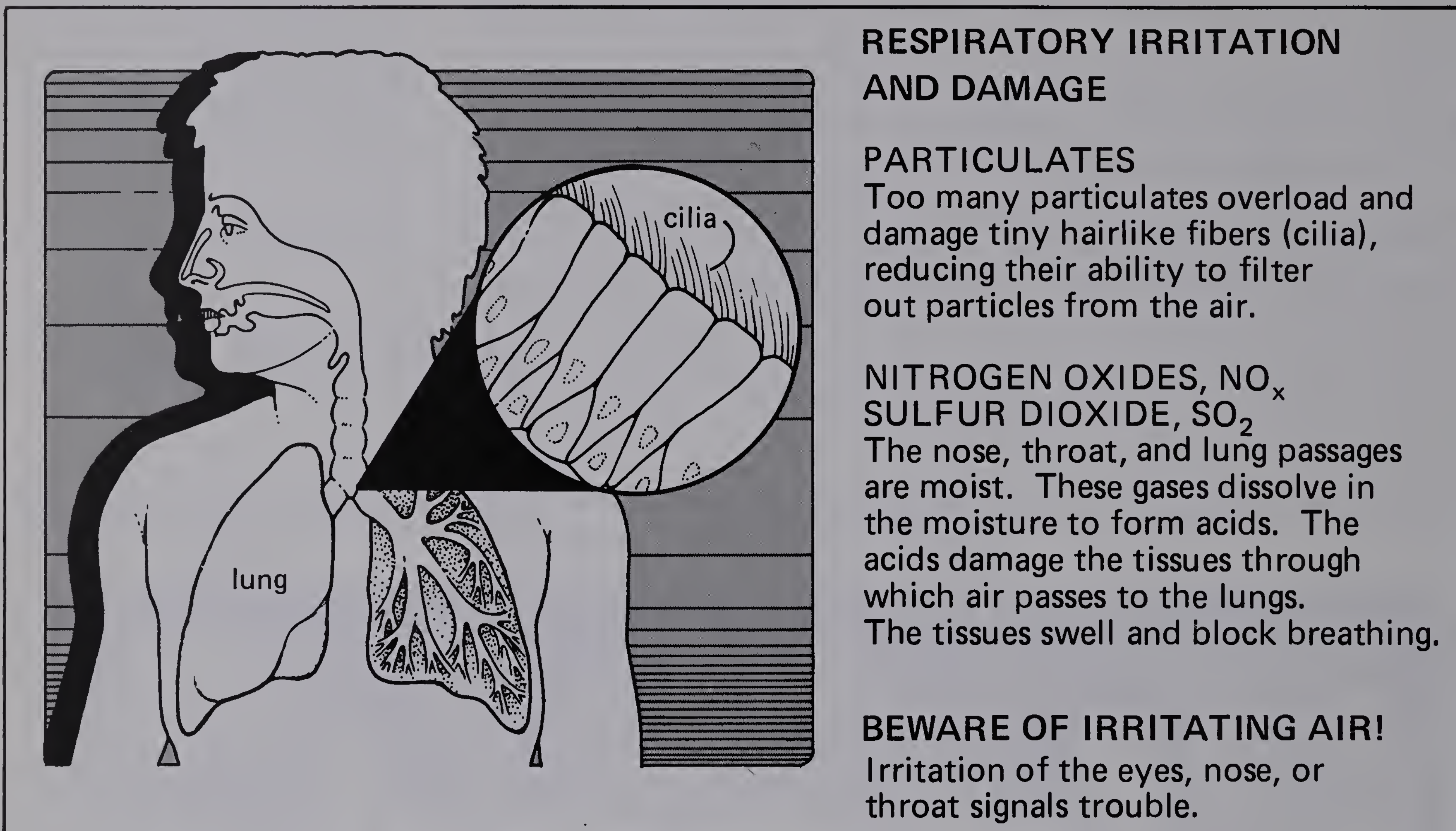


Figure 6-2

- 6-2. How do particulates, such as dust and soot, affect people?

- 6-3. Sulfur dioxide and nitrogen oxides can irritate the eyes and skin. What else can they do to the body?

6-2. They damage the cilia and interfere with their ability to filter out dust from the air.

6-3. They irritate the respiratory passages and react with moisture to produce acids.

By themselves, hydrocarbons don't irritate or poison much. They are dangerous only in very high concentrations. But combined with other pollutants from cars, they form a mixture that sunlight cooks into dangerous smog. This smog is worst on sunny summer days around noon. Look at Figure 6-3 below.

Auto smog is called *brown* or *photochemical* smog to distinguish it from gray or coal smog. Peroxyacetyl nitrate, PAN, is the most irritating photoproduct.

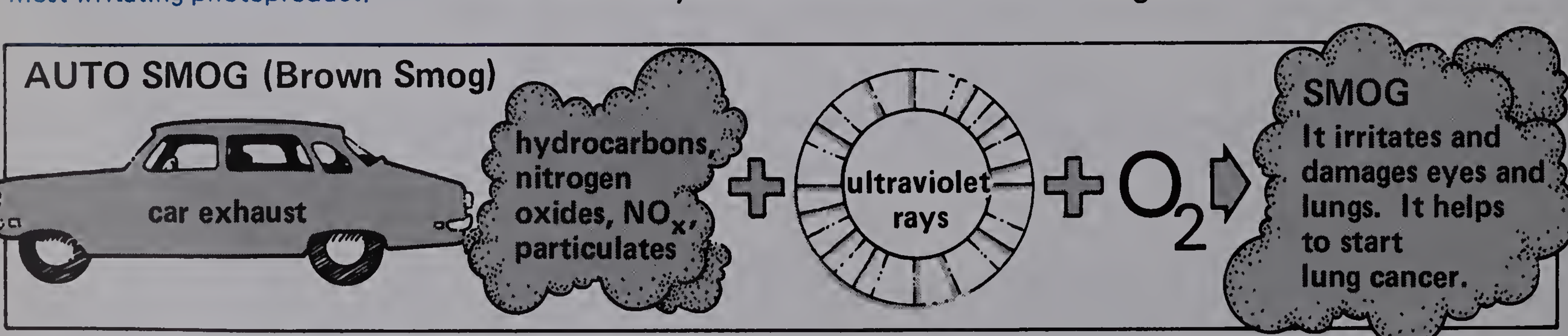


Figure 6-3

• 6-4. Hydrocarbons are dangerous only at high concentrations. How are they changed to new poisons that are dangerous at lower concentrations?

6-4. Hydrocarbons in sunlight combined with other pollutants and oxygen form smog.

☆ **6-5. What effect does auto smog have on people?**

6-5. It irritates and damages eyes and lungs and helps to start lung cancer.

A very similar occurrence produces a different kind of smog. Figure 6-4 below shows how smog forms from the burning of coal. This smog is worst on cold, foggy winter mornings.

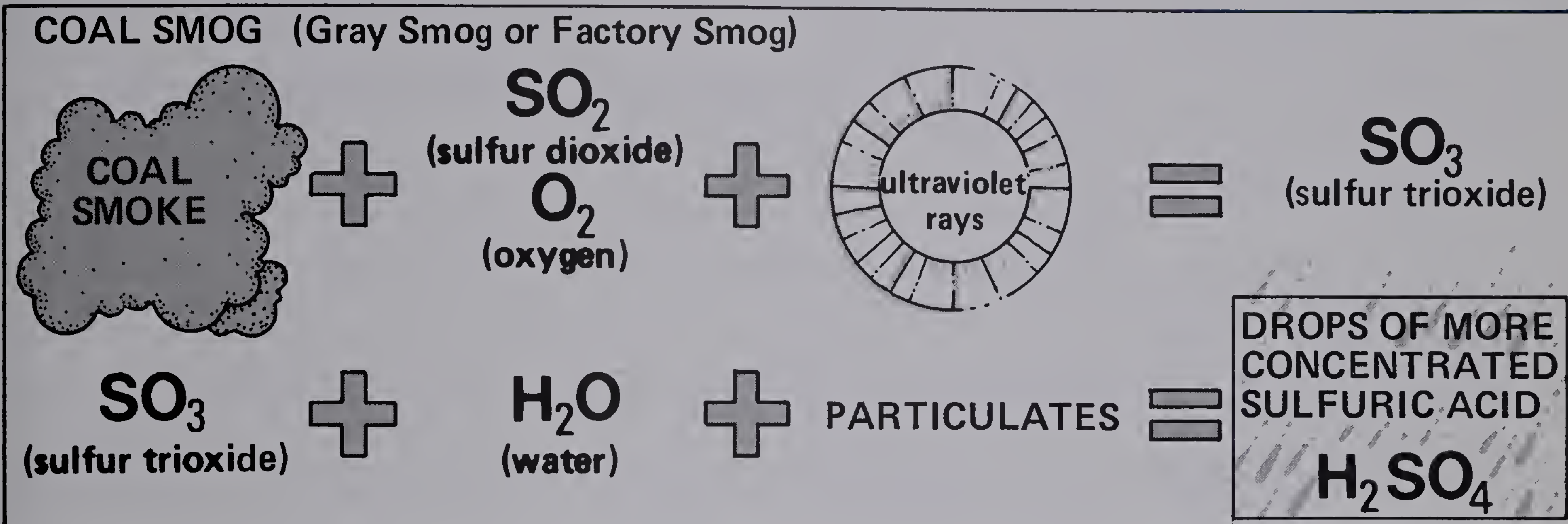


Figure 6-4

By itself, SO_2 dissolves in the moisture of nose, throat, and lung passages, forming irritating acid. But combined with smoke and fog, SO_2 produces more concentrated acid in droplet form. This more concentrated acid does much more damage.

☆ **6-6. What can make smog deadlier than the pollutants it is made from?**

6-6. SO_2 can combine with smog to form more concentrated acid, which does much more damage.

The effects of smog and the other air pollutants on plants, including trees, vary with the kind of pollutant and with the kind of plant. Some plants are more sensitive than others. Figure 6-5 below shows some of the effects of certain pollutants on plants.

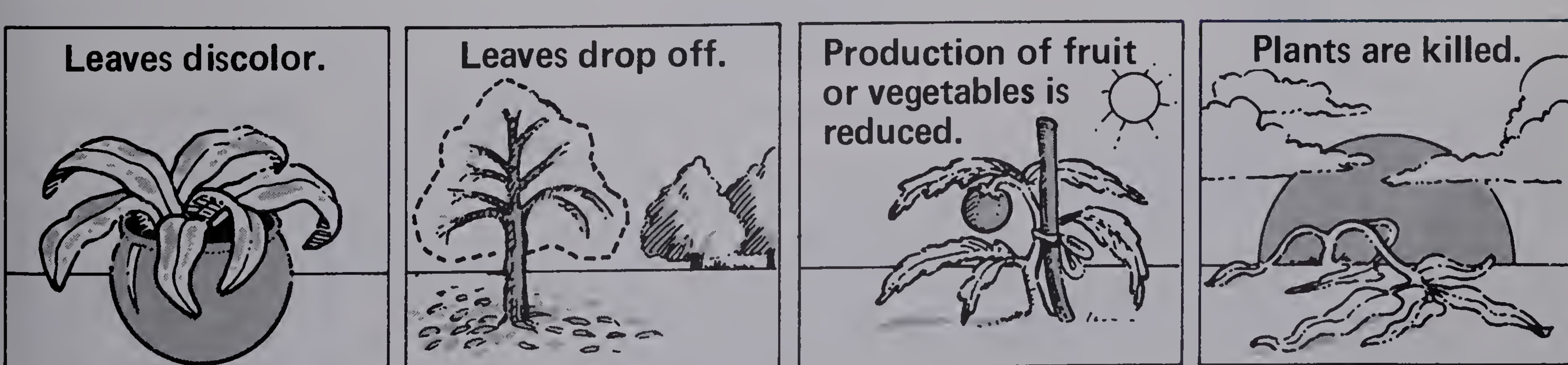


Figure 6-5

Pollutants in air enter a plant through tiny openings on the underside of the leaves. They can then kill the plant cells. When it rains or snows, smog, SO_2, and NO_x pollutants are washed down into the soil. There they form acids – and most plants don't grow well in highly acid soils.

6-7. A1, B1, C1, D2, E1

6-8. Smog can enter plant leaves and kill cells and it can enter soil, hurting growth. Damaged plants lose leaves, have decreased fruit yield, and may die.

☆ **6-7. Match each air pollutant with its effect on animals.**

Pollutant	Effect
A. Nitrogen oxides, NO_x	1. irritates and damages some part of the respiratory system
B. Sulfur dioxide, SO_2	2. interferes with the oxygen-carrying capacity of blood
C. Hydrocarbons	
D. Carbon monoxide, CO	
E. Particulates	

☆ **6-8. Describe two methods by which smog can damage plants. What effects does this damage have on plants?**

ACTIVITY EMPHASIS: As presented here, the oxygen–carbon dioxide cycle involves primarily plants, animals, microorganisms, and burning processes. Pure oxygen and carbon dioxide have several important uses in today's world.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter.

ACTIVITY 7: AIRY-GO-ROUND

The amounts and kinds of gases in air are fairly constant. But that doesn't mean things aren't always changing. Look at Figure 7-1 below.

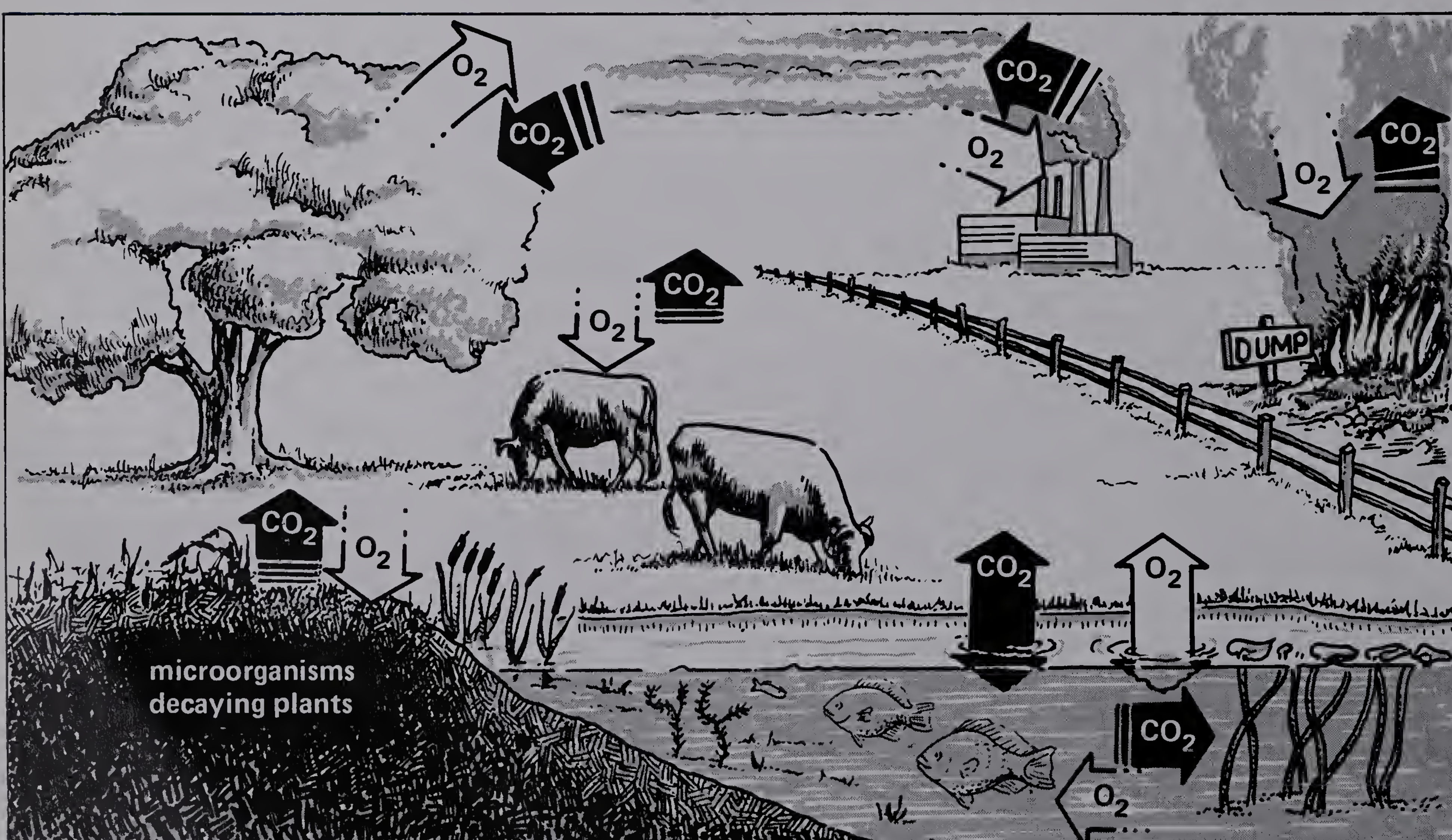


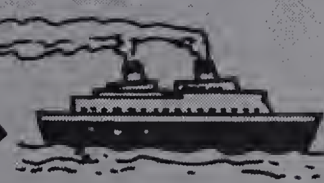
Figure 7-1

• 7-1. List all the things shown in Figure 7-1 (page 28) that take in oxygen, O_2, from the air.

7-1. The things that take in oxygen from the air are the cows, fish, microorganisms, the burning dump, and the power plant [and the pond water].

All the animals in Figure 7-1 are oxygen users. They take in oxygen, O_2, and give off carbon dioxide, CO_2. Animals use the oxygen to get energy from food. The oxygen-using process, which takes place in animal cells, is called *respiration*. It's very much like the burning process.

BURNING FUEL:

OXYGEN + FUEL →  + **CARBON DIOXIDE** + WATER

taken in — ENERGY — produced

RESPIRATION:

OXYGEN + FOOD →  + **CARBON DIOXIDE** + WATER

taken in — ENERGY — produced

• 7-2. Why must animals take in oxygen?

7-2. To get energy from foods

Decay microorganisms also take in O_2 and give off CO_2 by means of respiration. This helps decompose dead materials, such as leaves.

micro—very tiny
organism—living thing

☆ **7-3. What do respiration and burning take out of the air? What do they put into the air?**

7-3. They take oxygen out of the air and put carbon dioxide into the air.

• 7-4. Which things shown in Figure 7-1 (page 28) take in carbon dioxide?

7-4. The things that take in CO_2 are the water plants and the tree [and the water].

The living things that take in carbon dioxide, CO_2, are all plants. Plants must have CO_2, along with other substances, to make food, leaves, stems, and other parts. This process is called *photosynthesis* [fo-to-SIN-the-sis]. It requires light (that's the *photo* part), normally from the sun.

PHOTOSYNTHESIS:

ENERGY(sunlight) + **CARBON DIOXIDE** + WATER → FOOD + **OXYGEN**

taken in — produced

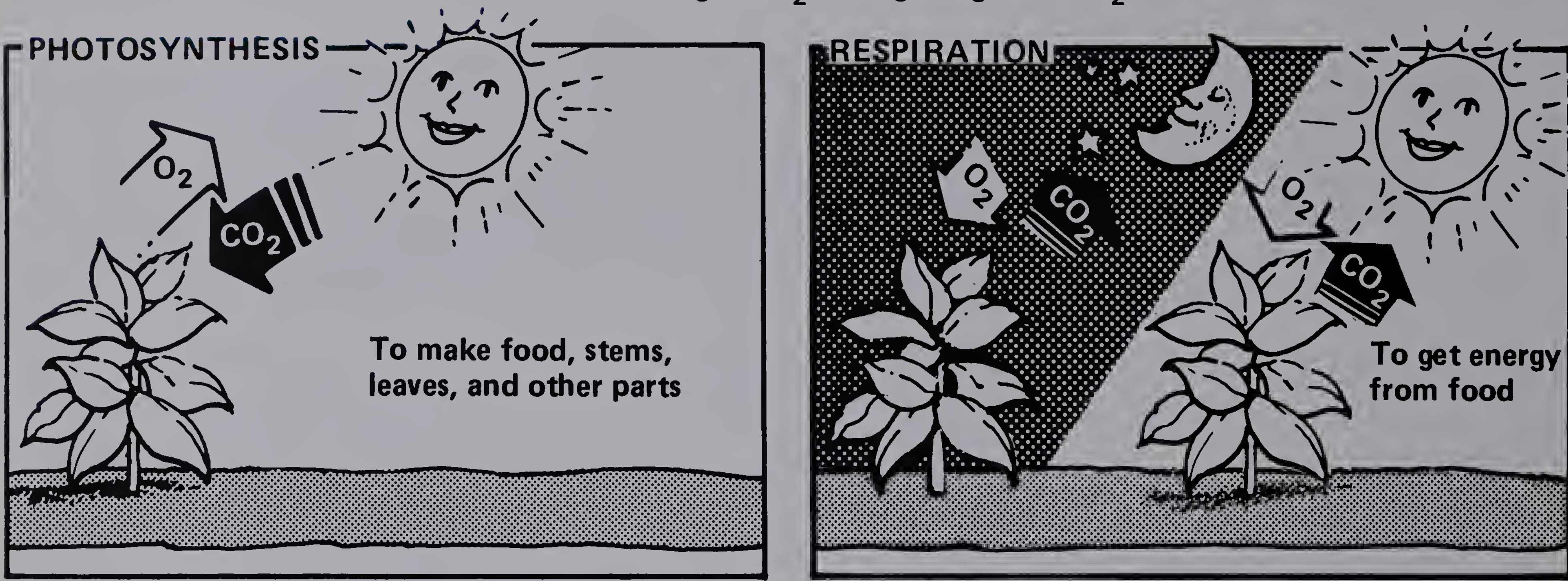
7-5. To make food, leaves, stems, and other parts by photosynthesis

7-6. Oxygen, O_2

- 7-5. Why do plants need carbon dioxide?

- 7-6. What does photosynthesis put back into the air?

By photosynthesis, plants make food – taking in CO_2 and giving up O_2. To get energy, the plants also carry on respiration – taking in O_2 and giving off CO_2.



Plants and animals seem to have a very good arrangement. Plants give off the oxygen that animals need. And animals give off the carbon dioxide that plants need.

The transfer of oxygen and carbon dioxide between living things and the air is called the *oxygen–carbon dioxide cycle.* The term *cycle* is used because the gases keep moving to and from the air over and over again. That's why the amounts of these gases in the air are always about the same.

7-7. Respiration (in both plants and animals) takes out O_2 and puts in CO_2. Photosynthesis in plants takes out CO_2 and puts in O_2 to keep the atmosphere as it is.

☆ **7-7. How do plants and animals help keep the oxygen–carbon dioxide cycle balanced?**

Anything that uses or produces oxygen or carbon dioxide is part of the cycle. Look at Figure 7-2 below.

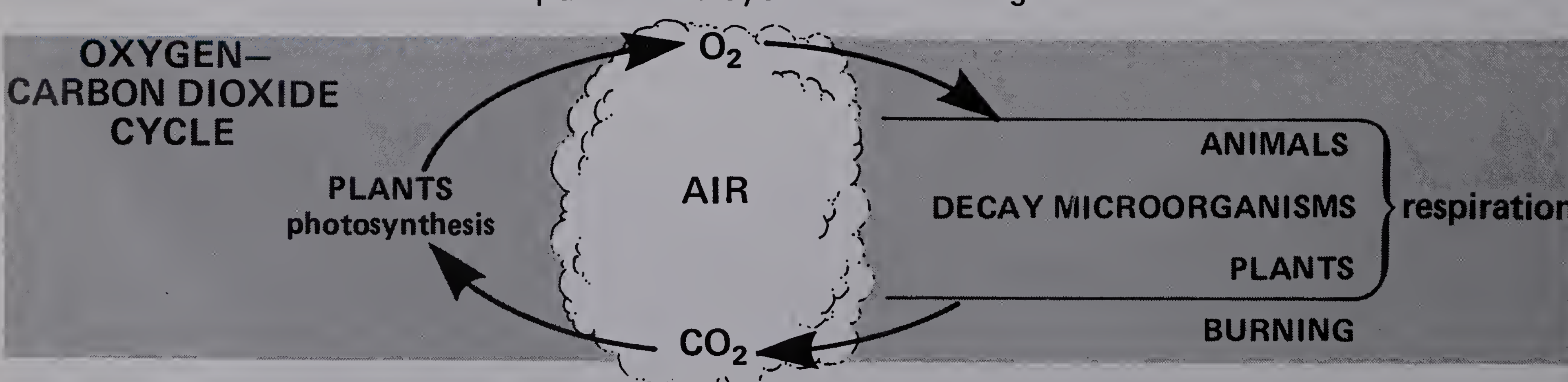


Figure 7-2

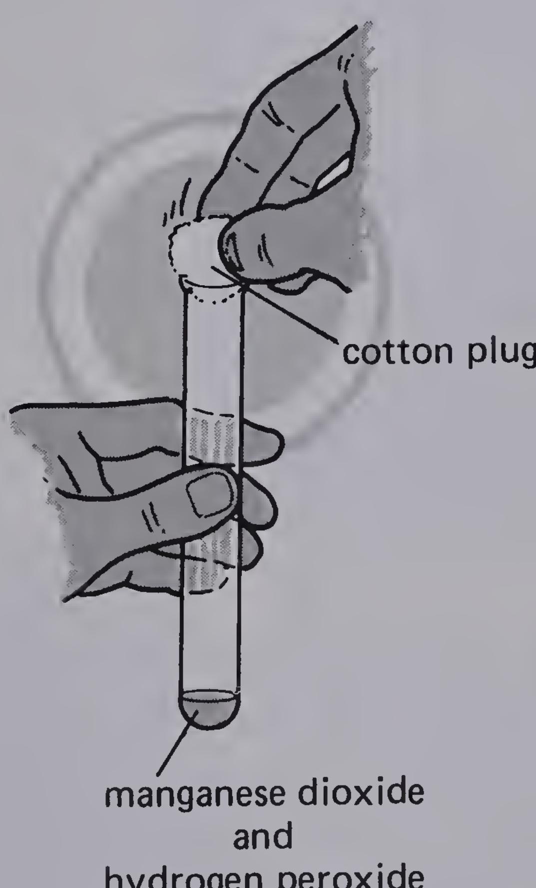
Suppose plants used more CO_2 and released more O_2 into the air. If more O_2 was in the air, what would happen? You can make oxygen-rich air and see. You'll need a partner and the following materials.

- safety goggles
- 10-ml graduated cylinder
- 2 medium test tubes
- cotton plug to make test-tube stopper
- 5 ml hydrogen peroxide, H_2O_2, 3% solution
- 0.1 g manganese dioxide, MnO_2
- wood splint
- safety matches

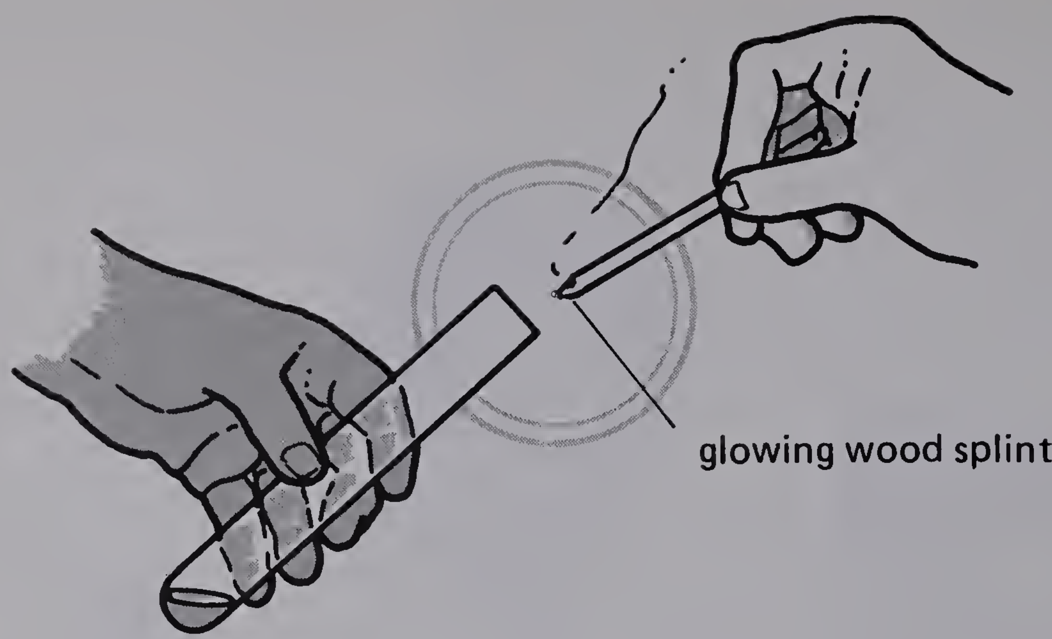
A. Put a few grains of manganese dioxide on the wood splint. Put the solid into one test tube by tapping the splint against the inside of the test tube.

Manganese dioxide catalyzes the reaction:
$2\ H_2O_2 \rightarrow 2\ H_2O + O_2$

B. Use the graduated cylinder to measure out 5 ml of hydrogen peroxide. Add it to the test tube, and quickly put the cotton plug into the opening. Oxygen is being produced by the reaction in the test tube.



C. Have your partner light the wood splint, and get it burning brightly. Then blow out the flame, leaving the splint glowing.



D. Remove the cotton plug, and have your partner quickly stick the glowing end of the splint into the test tube.

E. Repeat Steps C and D, using a clean test tube filled with just plain air.

7-8. In the oxygen-rich air, the splint burst into flame. In the plain air, it just continued to glow.

● 7-8. What happened to the glowing splint in the oxygen-rich air? In the plain air?

Sometimes faster or hotter burning is desirable. In those cases, pure oxygen or oxygen-rich air is used. Look at Figure 7-3 below. It shows some uses for oxygen.

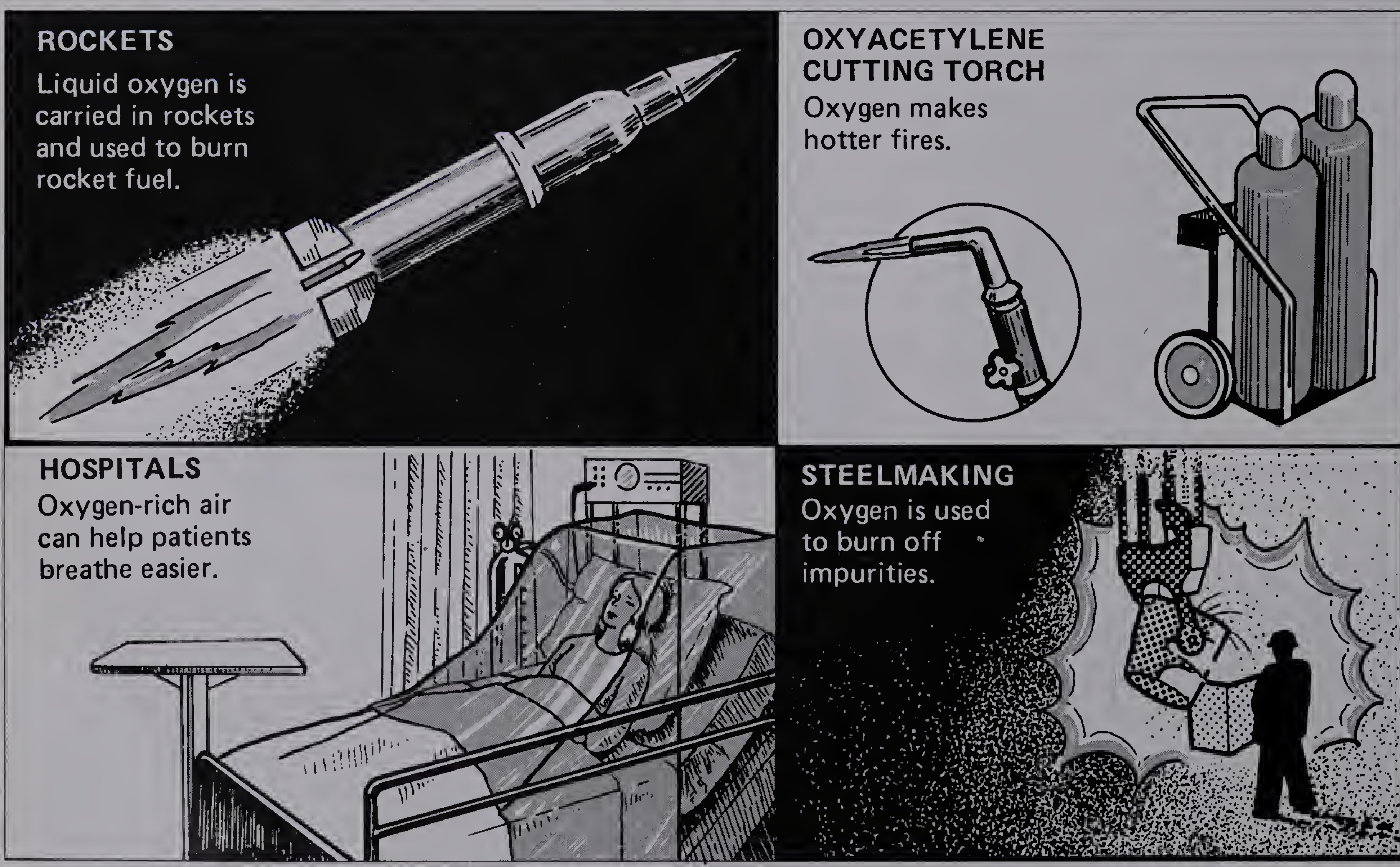


Figure 7-3

☆ **7-9. Name one use of oxygen-rich air and three uses of pure oxygen.**

7-9. Medical treatment; promoting burning of rocket fuels, acetylene torches, and burning off impurities in molten steel

Carbon dioxide, CO_2, gas also has important uses in its pure form. Figure 7-4 below shows some uses for pure CO_2.

7-10. To keep things cold; fire extinguishers and soda pop

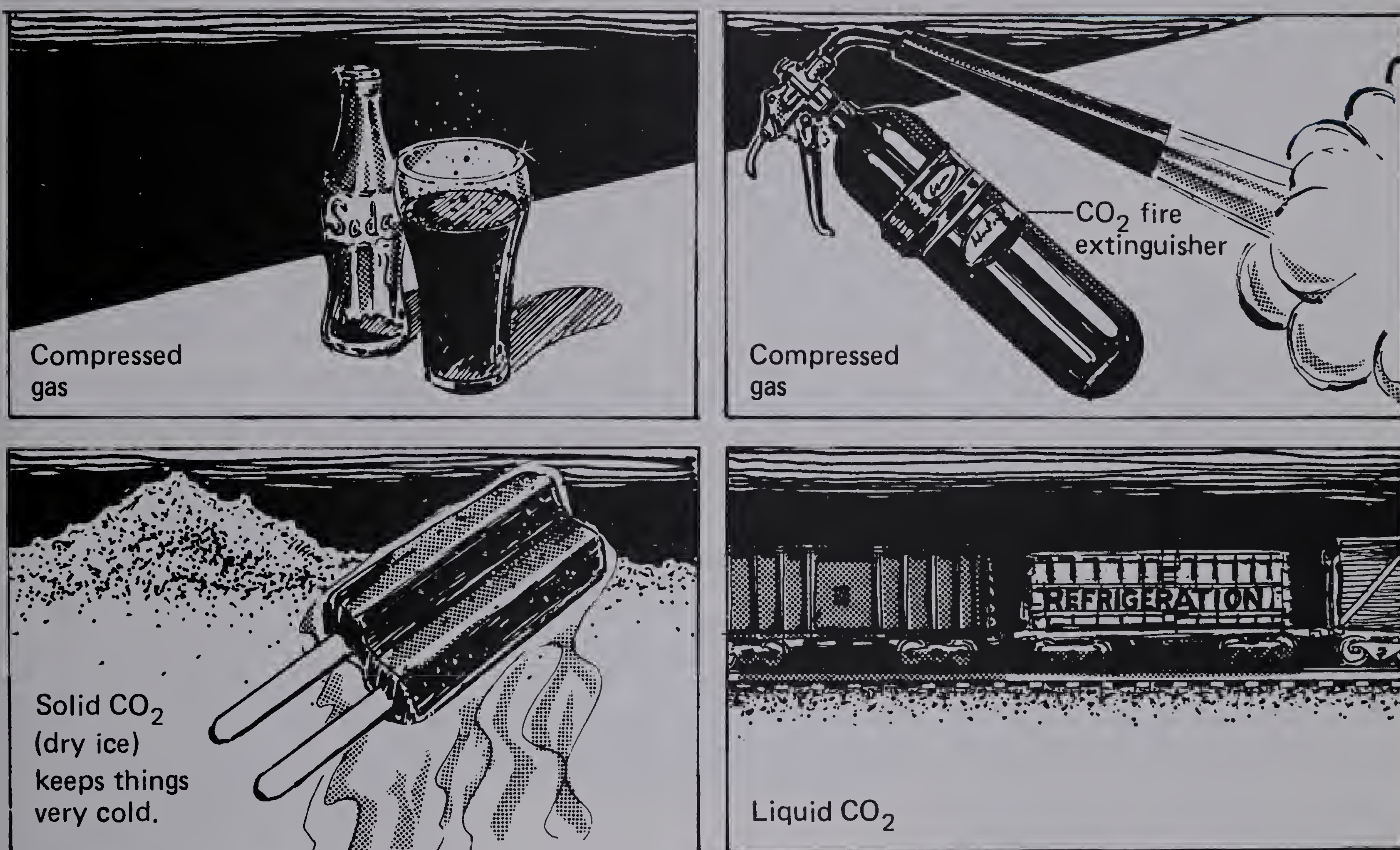


Figure 7-4

☆ **7-10. What are liquid and solid CO_2 used for? Gaseous CO_2?**

ACTIVITY 8: BASHFUL NITROGEN

Nitrogen is used to build proteins. Every single living cell in plants and animals contains some protein. So nitrogen is very important to all living things.

There's a lot of nitrogen in the air. But the nitrogen atoms in the air are stuck together in pairs. And most plants and animals can't use that form of nitrogen. They need individual nitrogen atoms.

The problem is how to change the nitrogen into a form plants can use to make proteins.

ACTIVITY EMPHASIS: Nitrogen is essential to living things, but they can use it only in combined form. Certain soil bacteria are the primary (terrestrial) nitrogen-fixing agents, converting atmospheric nitrogen to ammonium and nitrate compounds. The nitrogen cycle is completed by other bacteria that convert N compounds back to free N_2.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter.

Manufacturing processes have been developed to supplement natural fixation.

In the soil, certain very tiny organisms called *nitrogen-fixing bacteria* take nitrogen from the air. They *fix* it (turn it into a form that the plants can use) by a process called *nitrogen fixation*. Look at Figure 8-1 below.

NITROGEN FIXATION

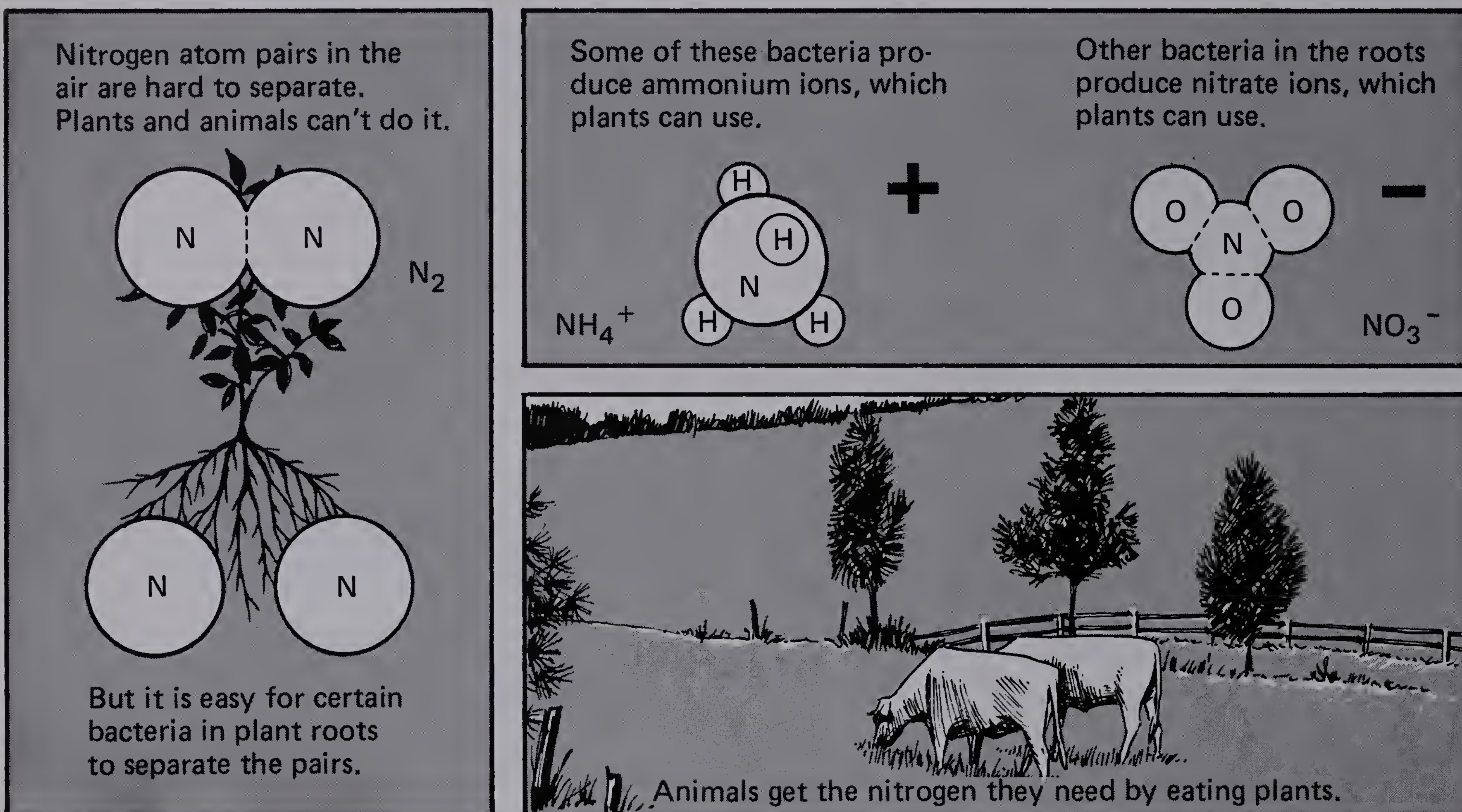


Figure 8-1

Bacteria of the genus *Rhizobium*, which form nodules in leguminous plant roots, fix nitrogen. Certain algae and free-living bacteria also fix nitrogen, as do lightning (producing nitrogen oxides) and certain photochemical processes. Nitrogen fixation by artificial processes — mainly for fertilizer production — is now a significant part of the total N_2 cycle.

In nitrogen fixation, the bacteria break the nitrogen atom pairs into single nitrogen atoms. The bacteria then attach the single nitrogen atoms to oxygen, O, or hydrogen, H, atoms. Plants *can* pull the nitrogen atoms away from O and H atoms.

The bacteria that do this live in the soil. Most actually live in the roots of plants like peas, clovers, and beans. The fixed nitrogen from the air is taken in by the plants through their roots.

These are called *nitrifying bacteria*.

• 8-1. Why can't plants and animals use nitrogen, N_2, directly from the air?

• 8-2. In what two forms can plants use nitrogen atoms? How do animals get nitrogen atoms?

8-1. N_2 is in the form of strongly bound pairs of N atoms. Plants and animals can't separate N_2 into the N atoms they need.

8-2. Plants can use nitrogen atoms as ammonium ions, NH_4^+, or nitrate ions, NO_3^-. Animals eat plants or other animals to get nitrogen atoms.

Lightning also fixes nitrogen, changing free nitrogen gas to nitrogen–oxygen compounds. So do some burning processes.

☆ **8-3. What removes nitrogen gas from the air and turns it into a form usable by plants? What's the process called?**

8-3. Bacteria in the soil, and lightning; *nitrogen fixation*

☆ **8-4. How does the form of nitrogen used by plants and animals differ from the nitrogen gas in the air?**

8-4. Nitrogen in the air is in the form of strongly bound pairs of nitrogen atoms. Usable nitrogen is fixed; that is, separate nitrogen atoms are combined with other atoms such as O and H.

Nitrogen fixation is always removing nitrogen gas from the air. Much of the nitrogen becomes part of plant and animal cells. This goes on constantly, but the air doesn't run out of nitrogen. There's a reason for that.

There is a process that puts nitrogen back into the air. When plants and animals die and decay, proteins and other nitrogen compounds are put into the soil. There, other kinds of bacteria turn the nitrogen in those compounds back into nitrogen gas. Look at Figure 8-2 below.

These bacteria are called *denitrifying bacteria*.

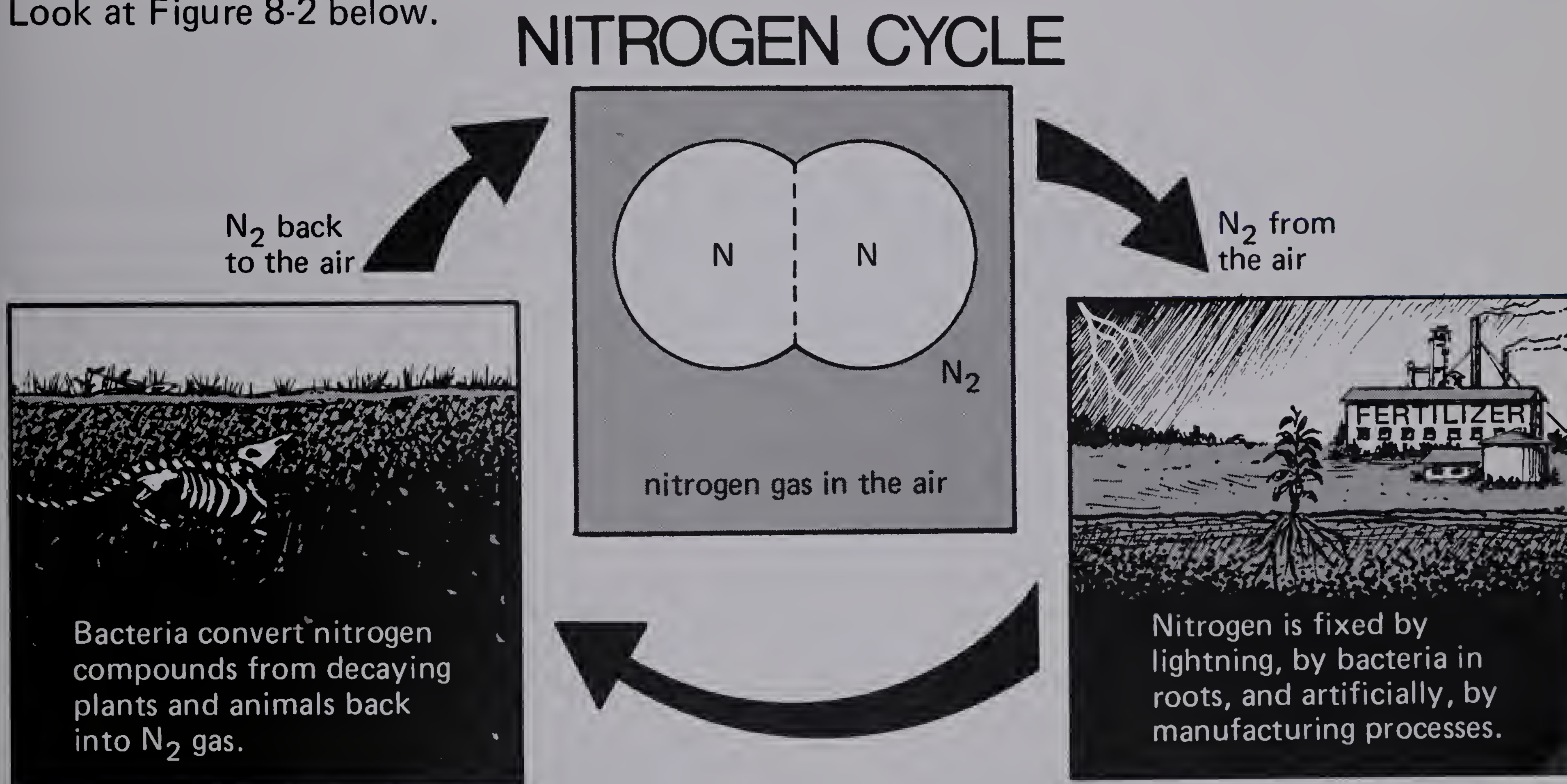


Figure 8-2

☆ **8-5. What returns nitrogen gas to the atmosphere?**

8-5. Certain bacteria in the soil convert nitrogen compounds in decaying matter to N_2.

Thus, nitrogen is removed from the air. It is fixed by bacteria or lightning into a form used by plants and animals. And it is returned to the air. This whole process is called the *nitrogen cycle*. It is called a *cycle* because nitrogen is always moving back and forth between the air and plant and animal life, over and over again.

☆ **8-6. What keeps about the same amount of nitrogen in the air all the time?**

8-6. The nitrogen cycle

The nitrogen cycle is one of the cycles that sustain life on Earth. Nitrogen atoms are needed by every animal and plant cell.

8-7. It sustains life on Earth by providing nitrogen needed by all living cells.

☆ 8-7. Why is the nitrogen cycle important?

To increase the amount of usable nitrogen for plants, ways have been found to artificially fix nitrogen gas — change it into usable forms. In these processes, nitrogen from the air is made into nitrogen compounds for use in fertilizers. The fertilizers help increase food production.

The first step in the artificial processes for changing nitrogen gas into useful compounds is usually to make nitrogen into ammonia, NH_3. This is done by combining the nitrogen with hydrogen.

NITROGEN	+	HYDROGEN	→	AMMONIA
N_2	+	$3\ H_2$	→	$2\ NH_3$
(from air)		(from petroleum)		

You can carry out one of the steps in making useful nitrogen compounds. You'll need the following materials.

Heavier copper wire, such as 18 gauge, works better. Fine wire cools too quickly.

Concentrated aqueous ammonia (28%) may be (improperly) labeled *ammonium hydroxide*.

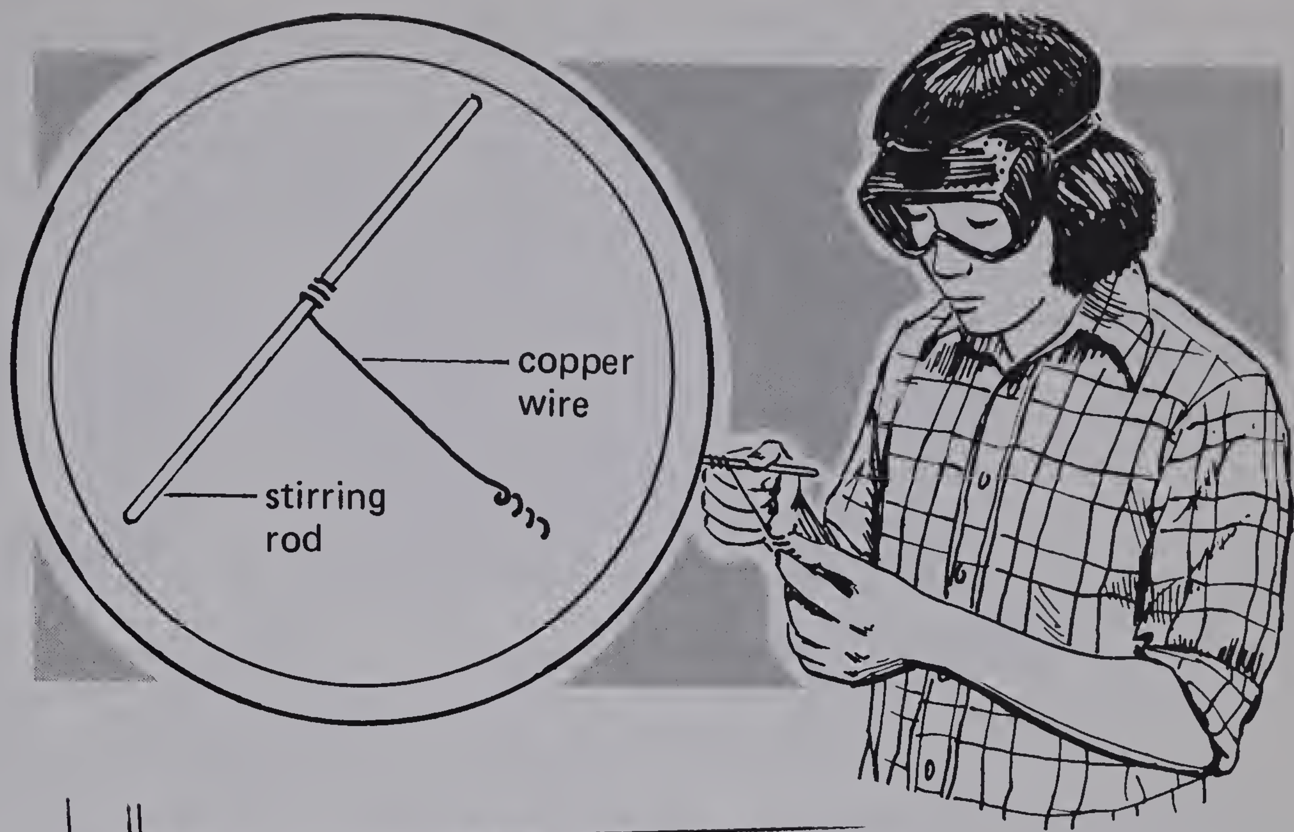
- safety goggles
- Bunsen burner
- 20 cm copper wire, clean and bright
- 10 ml concentrated ammonia water, $NH_{3(aq)}$
- 50-ml graduated cylinder
- stirring rod
- 250-ml Erlenmeyer flask

Have students work in a fume hood or in a well-ventilated area.

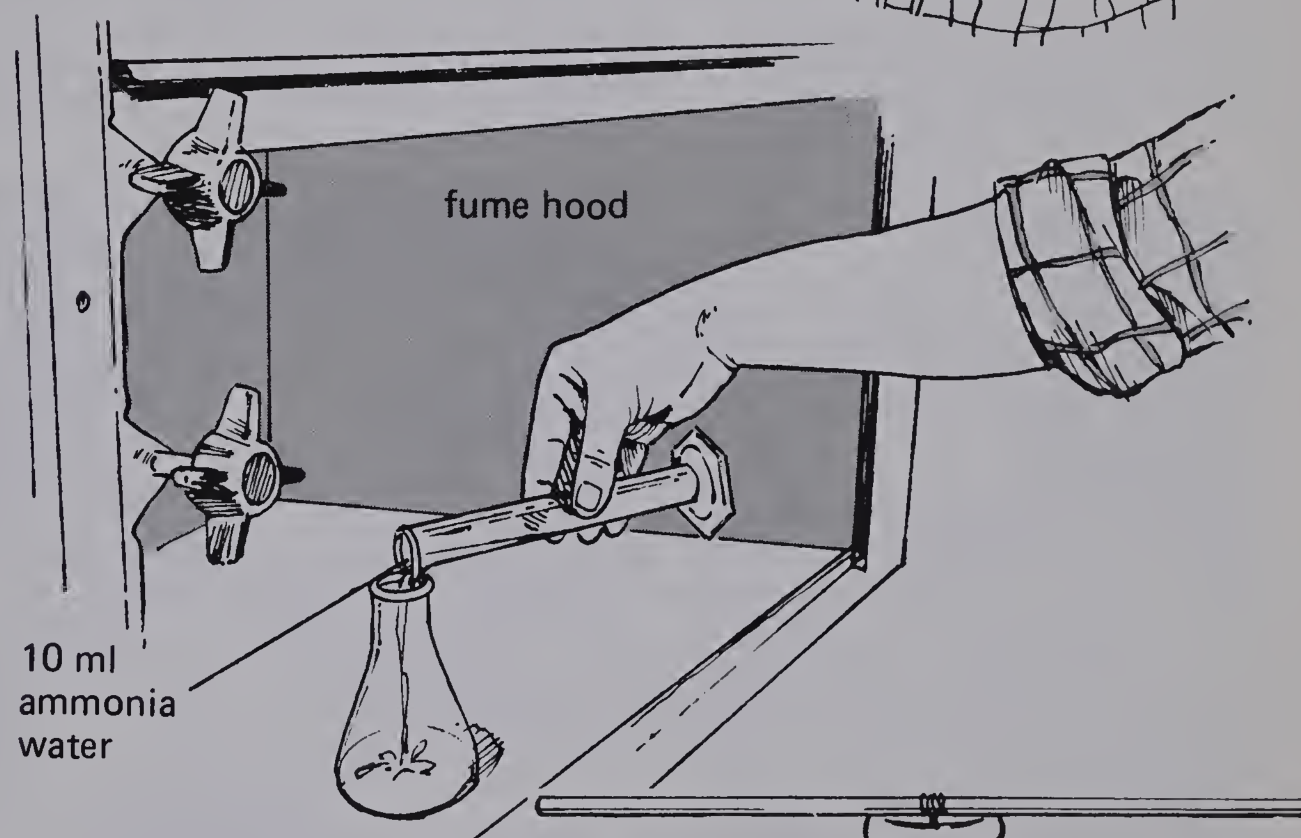
CAUTION Work in a fume hood or a well-ventilated area. Be very careful not to breathe the ammonia fumes. Avoid spilling the ammonia water on yourself or your clothes. It can damage your skin and clothes. If you spill any, wipe it up with a wet (not just damp) cloth or sponge. Tell your teacher.

If you need help using a burner or measuring volume, read "Resource Unit 17: Using a Burner" or "Resource Unit 5: Measuring Volume" now. Then begin Step A.

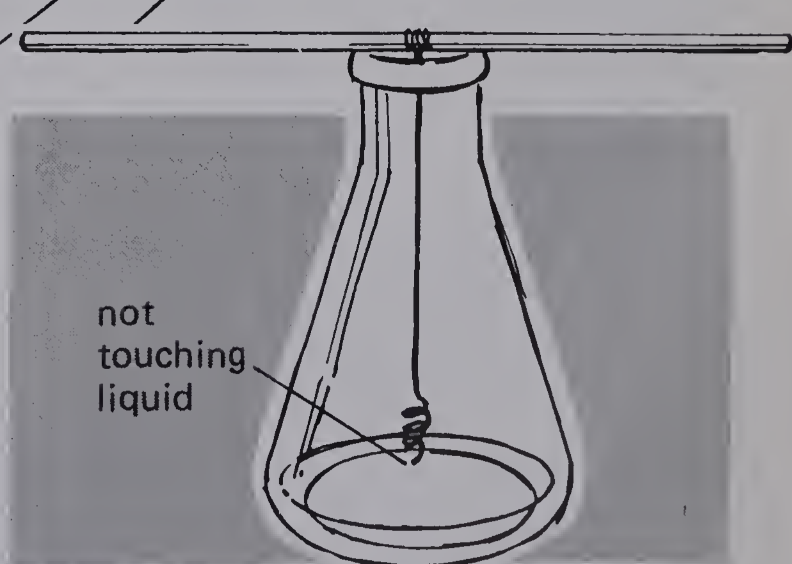
A. Wrap several turns of one end of the copper wire around the middle of the stirring rod. Twist the other end of the copper wire into a spiral as shown.



B. Carefully pour 10 ml of concentrated ammonia water into the flask. Immediately rinse the graduated cylinder with water, and set it aside.

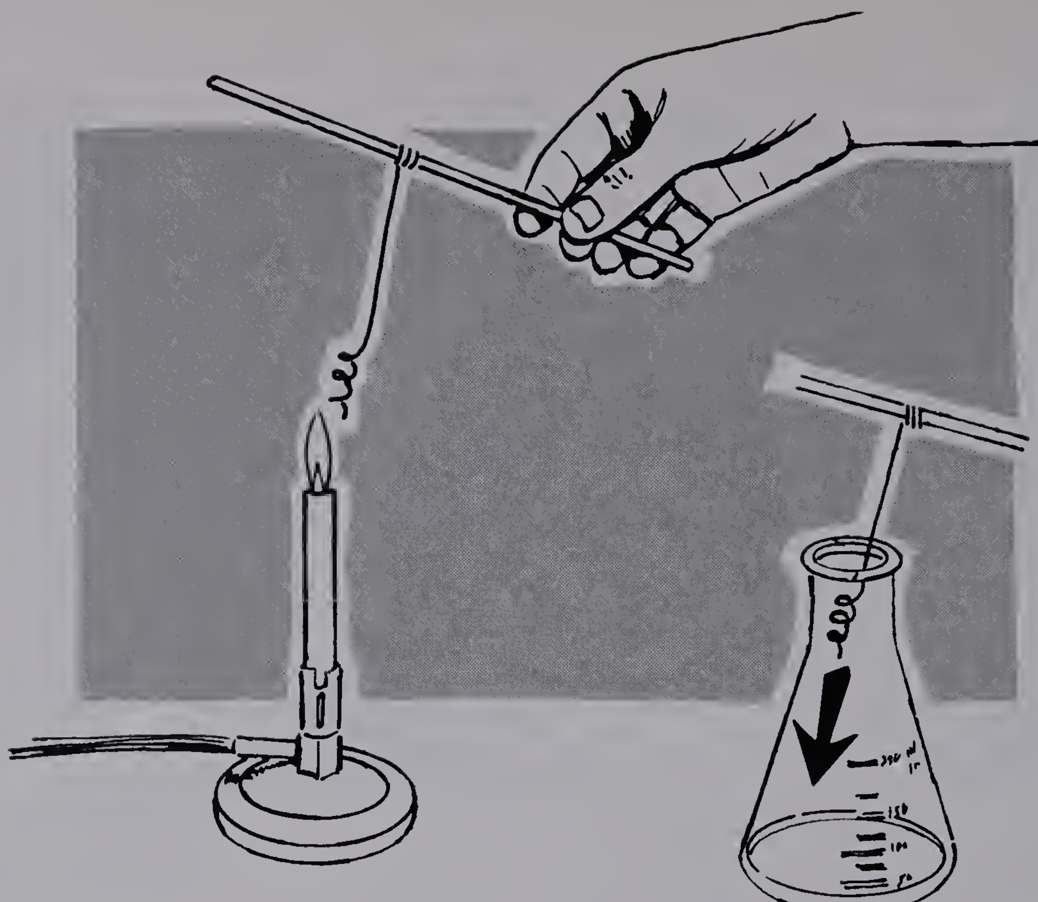


C. Place the copper wire assembly in the mouth of the flask to check the length of the wire. The wire should be close to, but not touching, the surface of the ammonia water. If necessary, adjust the length of the wire.



CAUTION	Touch only the stirring rod as you heat the wire. Don't touch any part of the wire. It can burn your hand.

If the student treats the copper wire too long, it may melt into two pieces. That is less likely to happen with thick wire.



D. Light the burner. Remove the stirring rod with the copper wire from the flask. Heat the wire until it glows very brightly, but don't touch it.

E. Immediately put the glowing end of the wire into the flask.

8-8. It continues to glow brightly.

The hot wire initiates and catalyzes the reaction: $4\ NH_3 + 5\ O_2 \rightarrow 4\ NO + 6\ H_2O$. Heat promotes the reaction $2\ NO + O_2 \rightarrow 2\ NO_2$. The visible fumes are NH_4NO_3, NH_4NO_2, and spalled copper oxides.

- 8-8. What happens to the copper wire in the flask?

When the ammonia reacted with oxygen from the air, energy was released. This energy made the wire continue to glow. The reaction also formed nitrogen oxides, NO_x.

One process of artificial nitrogen fixation uses electric arcs to make NO_x — imitating lightning. These nitrogen oxides can then be processed to make fertilizer.

Fixed nitrogen is used to make many different products. Figure 8-3 below shows some of them.

Figure 8-3

The nitrogen gas in air must be fixed before plants can use it. But there are some ways nitrogen can be used directly. First it's separated from the other gases in air. Then the pure nitrogen gas is used to protect molten metals from corrosion by the oxygen in air. Nitrogen in liquid form is used in refrigeration.

☆ **8-9. Name one use for liquid nitrogen, one for nitrogen gas, and four for fixed nitrogen.**

8-9. Refrigeration; protecting metals from corrosion by oxygen in air; making explosives, fertilizer, nitric acid, and ammonia water

ACTIVITY 9: TRADING CLEAN AIR

ACTIVITY EMPHASIS: There have to be tradeoffs between increased production and increased air pollution. It's not easy to decide what's best.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter.

People want the best food, clothing, shelter, medical care, transportation, and overall standard of living that they can get.

Producing things that people need or want and disposing of things they no longer want pollutes the air. Figure 9-1 below shows some of the pollution involved in the use of just one common metal, copper.

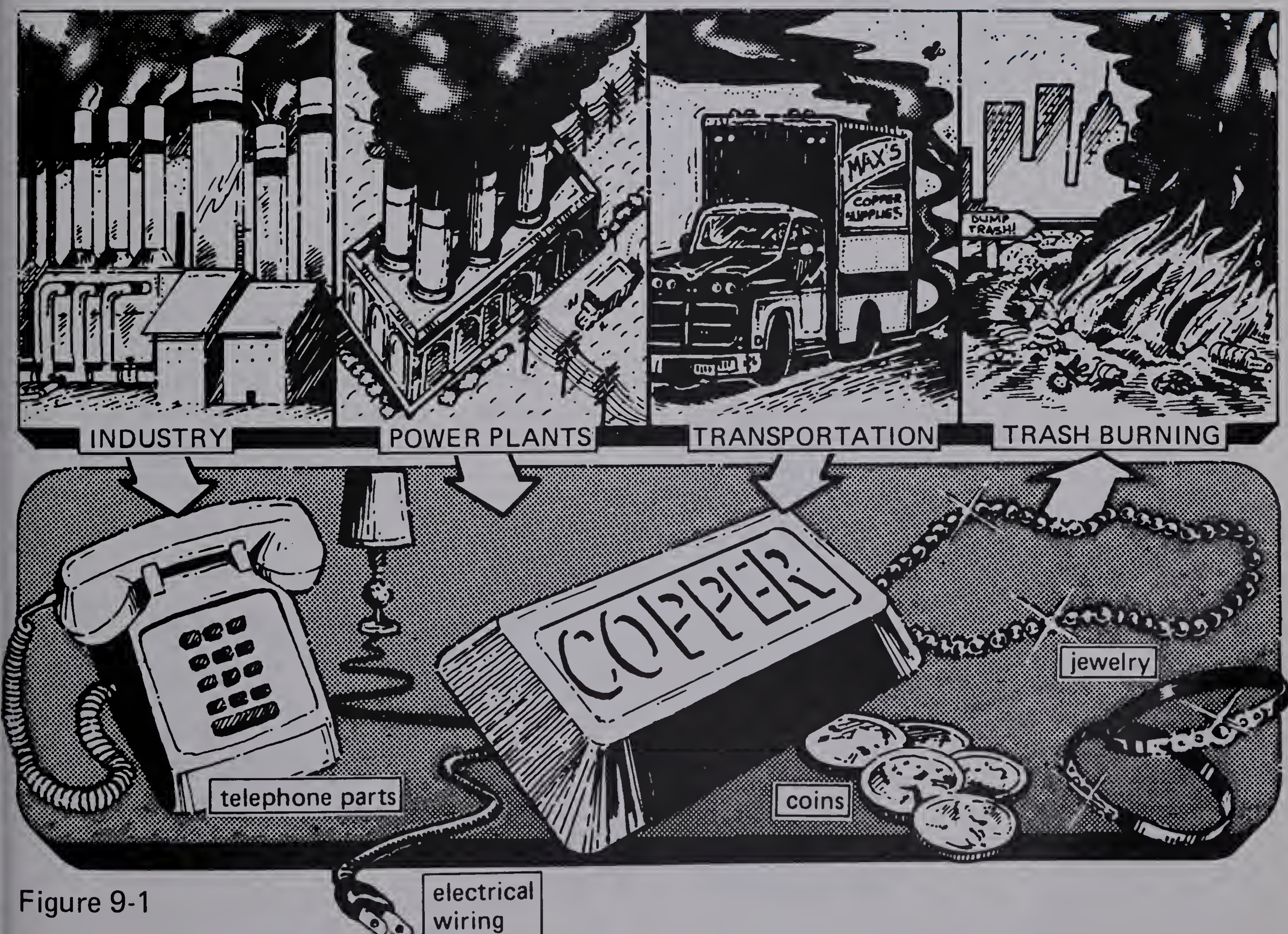


Figure 9-1

9-1. Industry makes the equipment for mining copper ore, extracting the metal, making it into products, and wrapping the products. Transportation is used for the ore, metal, products, and all the people who make them. Used products and packaging are disposed of as trash. All these activities need power supplied by power plants.

A short length of glass tubing with the ends fire-polished can be used instead of the medicine dropper. However, the dropper is easier to insert and remove from the stopper. The investigation can also be done without the glass chimney or even the stopper, but there's usually more spread of SO_2 and less success with the litmus test.

• 9-1. How is each source of air pollution shown in Figure 9-1 (page 39) connected to the making, use, and disposal of copper and its products?

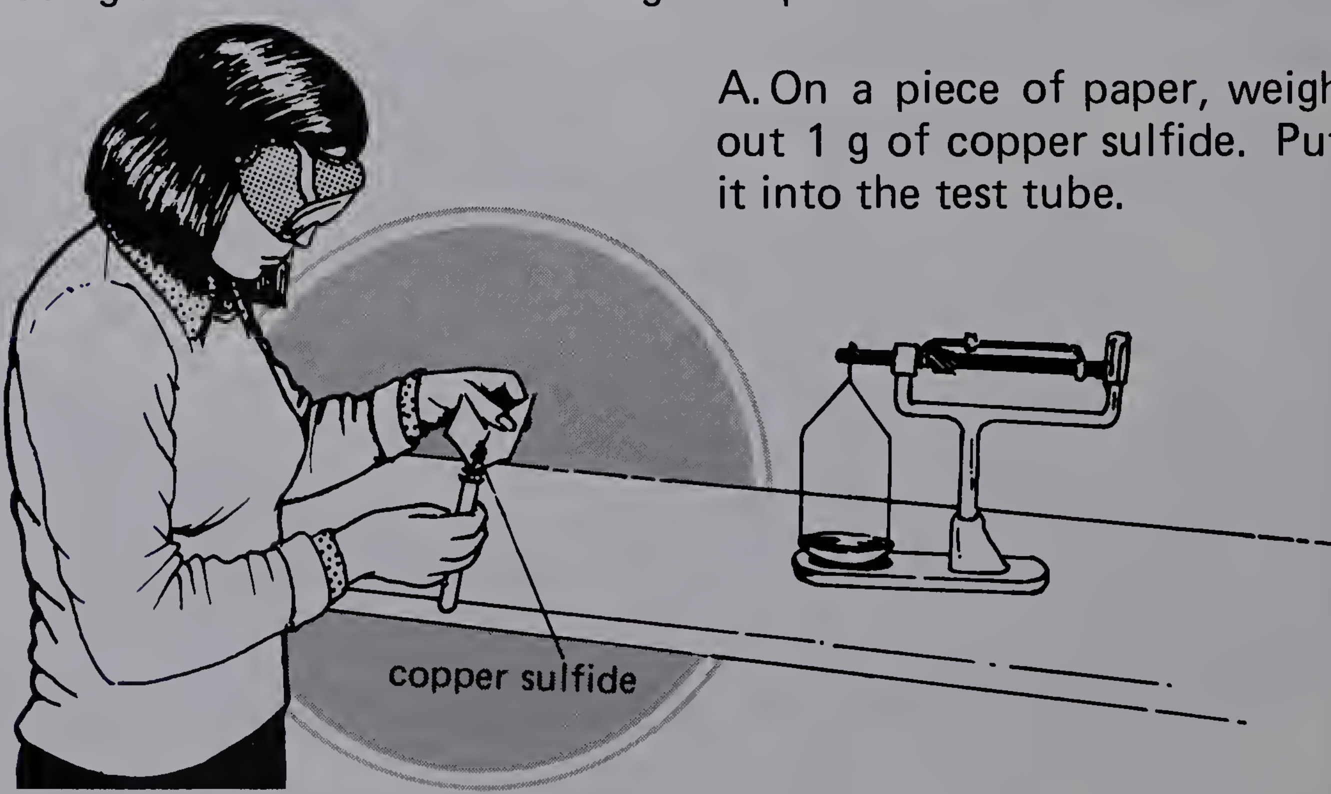
In nature, most copper is combined with other materials in *ores*. Air pollution results from mining copper ores and getting the copper from the ore. It also results from the use of copper in many products.

Air pollution in the actual freeing of copper from its ores is only one small part of the overall pollution that the use of copper produces. But it's a part that counts. You can get an idea of part of the price people pay in air pollution by examining one way of getting pure copper metal. The most common copper ore is copper sulfide — copper combined with sulfur. To get pure copper metal, you have to separate the sulfur from the copper sulfide. Try it. You'll need a partner and these materials.

safety goggles
1 g copper(II) sulfide, CuS
medium test tube
one-hole rubber stopper to fit test tube
medicine dropper
Bunsen burner
blue litmus paper strip
test-tube holder
sheet of paper
balance

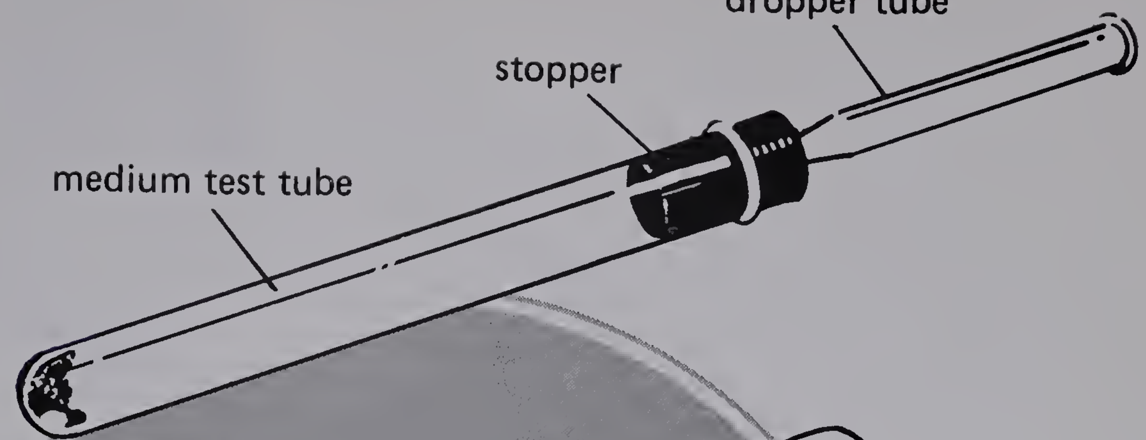
If you don't know how to use a balance or a burner, read "Resource Unit 10: Using Balances" or "Resource Unit 17: Using a Burner" now. Then begin Step A.

A. On a piece of paper, weigh out 1 g of copper sulfide. Put it into the test tube.

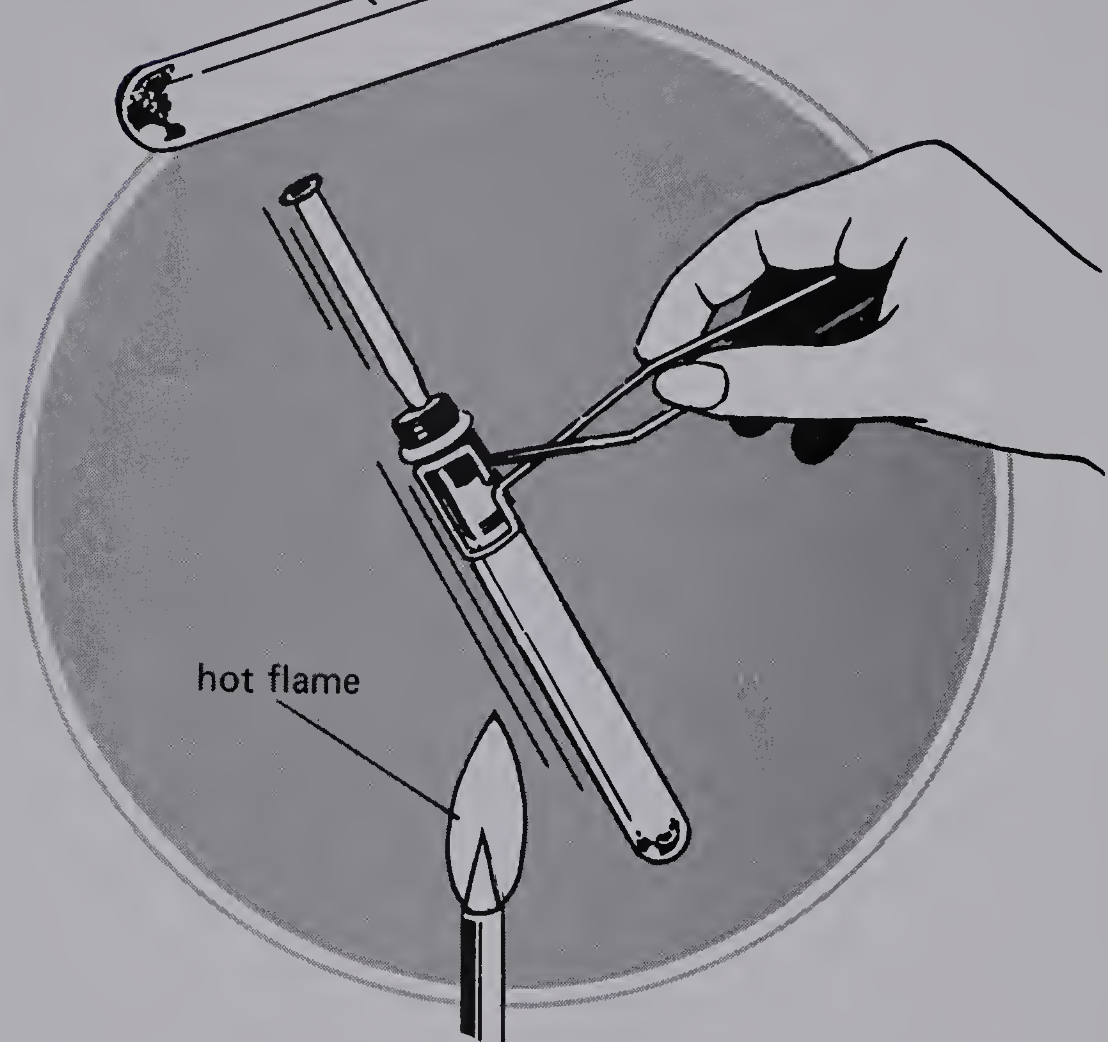


Caution students not to force the dropper tube into the stopper.

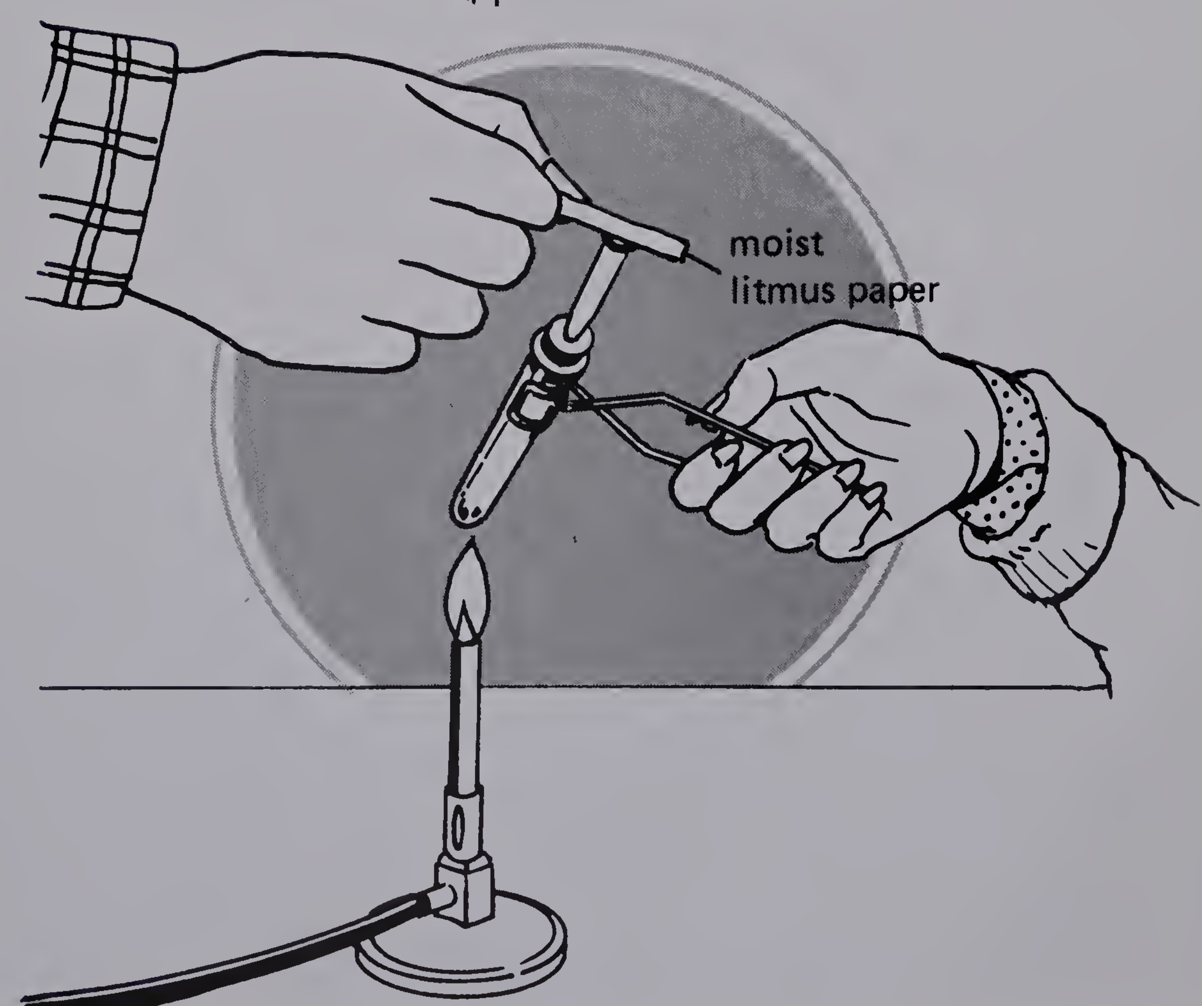
B. Remove the medicine-dropper bulb, and fit the tapered end of the dropper tube into the rubber stopper. Put the stopper in the end of the test tube.



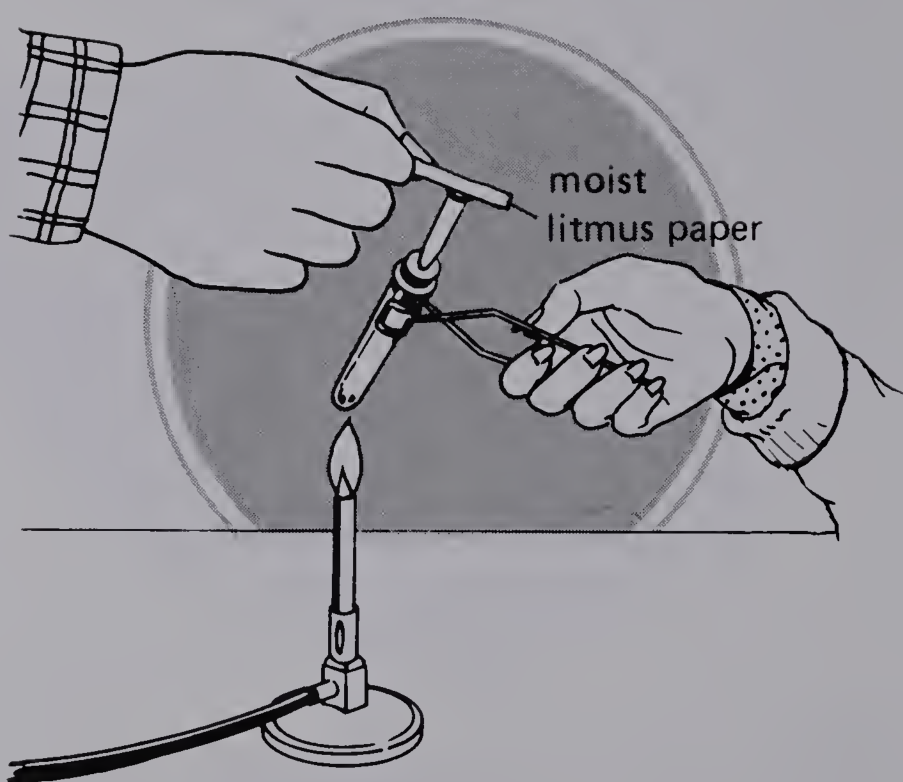
C. Light the burner, and adjust it for the hottest flame. Heat the test tube, gently at first, just above the copper sulfide. Move it back and forth over the flame. Then, slowly move the flame down to the bottom of the test tube.



D. Have your partner moisten the piece of blue litmus paper with water. Then, have your partner hold the litmus paper partly across the opening of the dropper tube. Continue heating the test tube.



E. Have your partner remove the litmus paper. Then have your partner carefully smell any gases being given off, as shown. Keep your faces away from the dropper tube.



F. Have your partner hold the litmus paper again. Now, strongly heat the copper sulfide until it glows. Have your partner record any changes in the appearance of the test tube and the litmus paper. Turn off the burner. Let the test tube cool.

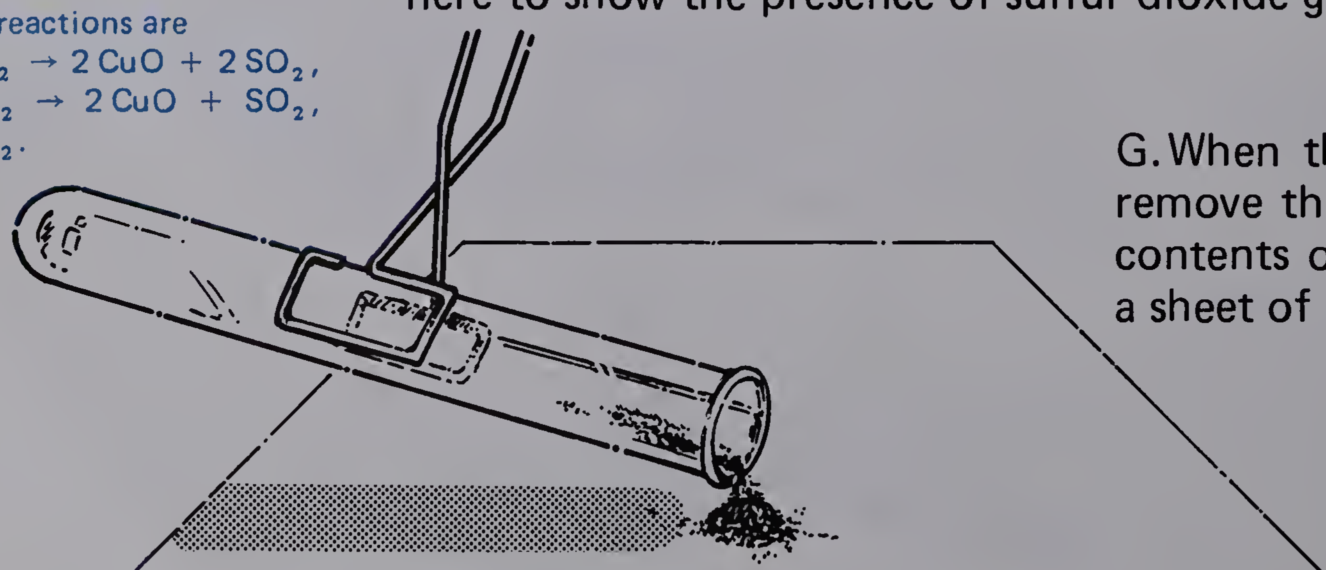
9-2. There was yellow sulfur deposited in the neck of the test tube.

9-3. The odor was sharp and disagreeable. The litmus paper turned red. Bleaching may have been observed.

The possible reactions are
$2\,CuS + 3\,O_2 \rightarrow 2\,CuO + 2\,SO_2$,
$Cu_2S + 2\,O_2 \rightarrow 2\,CuO + SO_2$,
$S + O_2 \rightarrow SO_2$.

● 9-2. What evidence was there as you heated the test tube that the copper and the sulfur in copper sulfide were being split apart? (Hint: Sulfur is yellow.)

● 9-3. Describe the gas that was produced — its odor and its effect on blue litmus. (Hint: Moist blue litmus paper was used here to show the presence of sulfur dioxide gas.)



G. When the test tube is cool, remove the stopper. Pour the contents of the test tube onto a sheet of paper.

● 9-4. Describe the appearance of the solid. How is it different from what you started with?

9-4. The dark green of CuS is replaced by the black of CuO [but the change is often difficult to see].

When copper sulfide is heated, some of the sulfur is removed. You saw it in the test tube. It was yellow. Some of the sulfur that was removed combined with oxygen in the air to form sulfur dioxide. Sulfur dioxide gas has a choking, unpleasant odor and turns moist blue litmus to red.

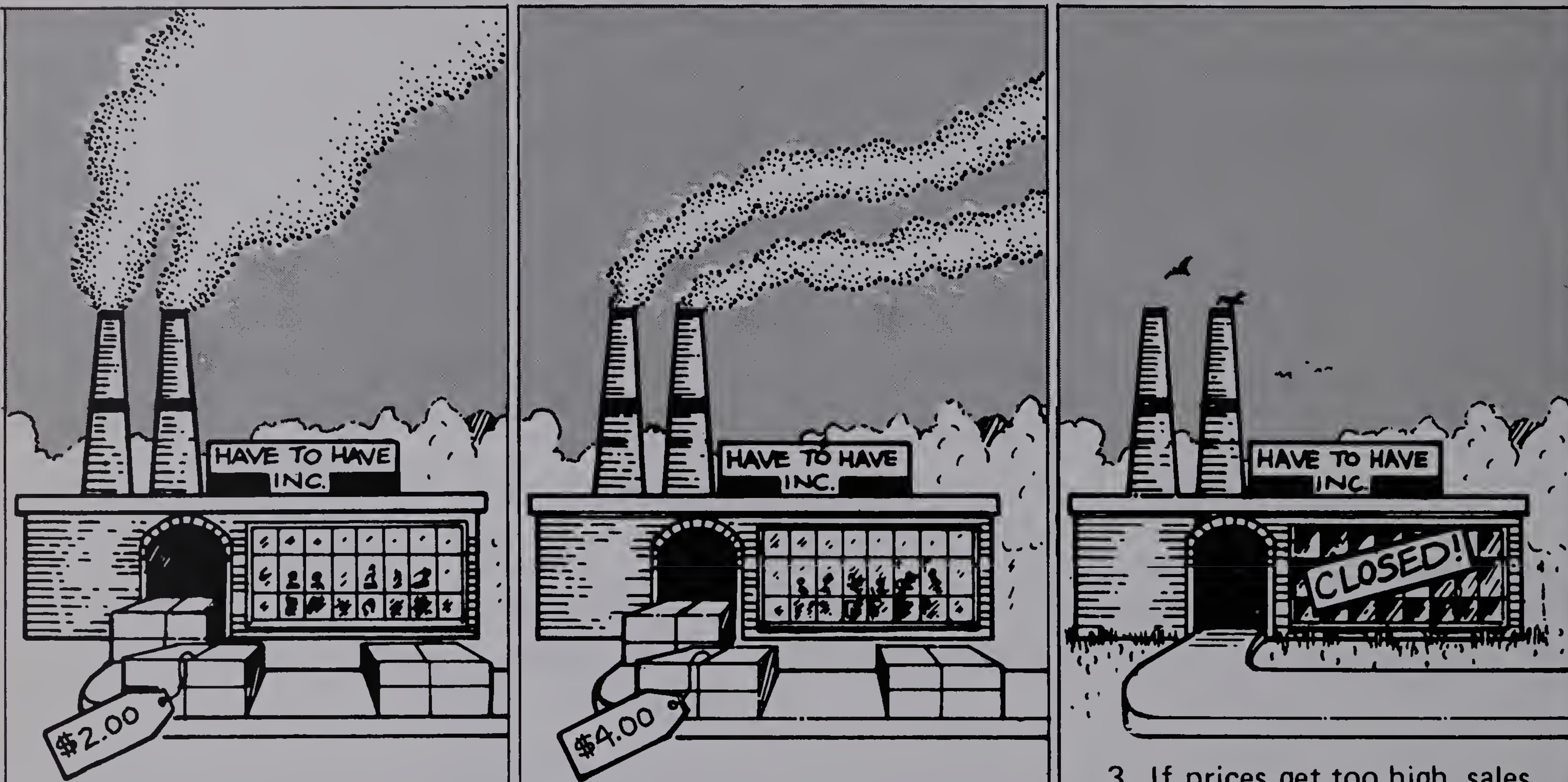
The initial heating produces some CuO and SO_2, as well as Cu and S. Further heating in the absence of air results in copper metal from this reaction: $CuS + 2CuO \rightarrow 3Cu + SO_2$. The net reaction is $CuS + O_2 \rightarrow Cu + SO_2$.

The black copper oxide and sulfide mixture left in the test tube has to be heated still further under airless conditions to get pure copper metal. And that step produces even more sulfur dioxide!

As you saw, in getting the very useful copper metal, you also produce the very unpleasant by-product sulfur dioxide gas. Unless that gas is trapped, it goes into the air and pollutes it. In the air, it is harmful to plants and animals.

By-products are formed along with nearly everything manufactured. Many of these, such as sulfur dioxide, are serious air pollutants. If the by-products aren't adequately removed, they add to industrial smog.

Perhaps people can have everything they want and still not allow the pollutants to escape into the air. But clean air is no longer free. Consider the following situation.



1. To keep pollutants out of the air, people must spend money on devices to trap them.

2. This makes the product being produced more expensive. Sometimes, it gets so expensive that people can no longer afford it.

3. If prices get too high, sales go down and the plant must close. The workers lose their jobs. And people don't get the products they wanted in the first place.

9-5. 1. People won't have fertilizer. Some people will lose their jobs. 2. Air pollution may harm plants and animals. 3. The cost of control devices would increase the price of the fertilizer, maybe so much that the customers can no longer afford to buy the fertilizer.

☆9-5. Suppose that a fertilizer plant in your community produces an air pollutant as a by-product. A citizens' group proposed three possible courses of action.
1. Close down the plant.
2. Fertilizer is needed, so allow the plant to continue.
3. Require the plant to buy expensive pollution-control devices.
Explain why none of these actions are completely satisfactory.

So people compromise. How much air pollution must people put up with to have the things they want and need? What kinds of things do people make that aren't worth what they cost in increased air pollution? There are no easy answers to problems like these.

Each case has to be considered separately. And people must decide what is the best compromise in each case.

9-6. The most satisfactory answer might be a compromise that balances the benefit of fertilizer production against the dangers of air pollution.

☆9-6. What might be the most satisfactory answer to the problem stated in Question 9-5 above?

ACTIVITY 10: IS AIR GETTING CLEANER?

ACTIVITY EMPHASIS: Recent trends in the amounts of various air pollutants are derived from graphs made by the student. The student decides which trends indicate successful control of air pollutants.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter.

Air pollution is everywhere. If you live in the city, you see it, smell it, taste it, and feel it. Away from large cities, you don't get as much. But it's there, too.

Air pollution affects the lungs, burns the eyes, and rots away nylon clothes. It corrodes the paint and steel on cars and ruins windshield-wiper blades. It damages crops, blackens the sky, and dirties everything. It ruins people's health, and it even kills.

To fight air pollution, you have to know how fast it's growing. For each pollutant, you have to know what the *trend* is — the general change over time. One of the best ways to see trends is with graphs. Look at the graph in Figure 10-1 below.

During the first seventy years of this century and even before that, air-pollution levels increased. Cars were first manufactured in the early 1900s. As the number of cars increased, so did the pollution levels.

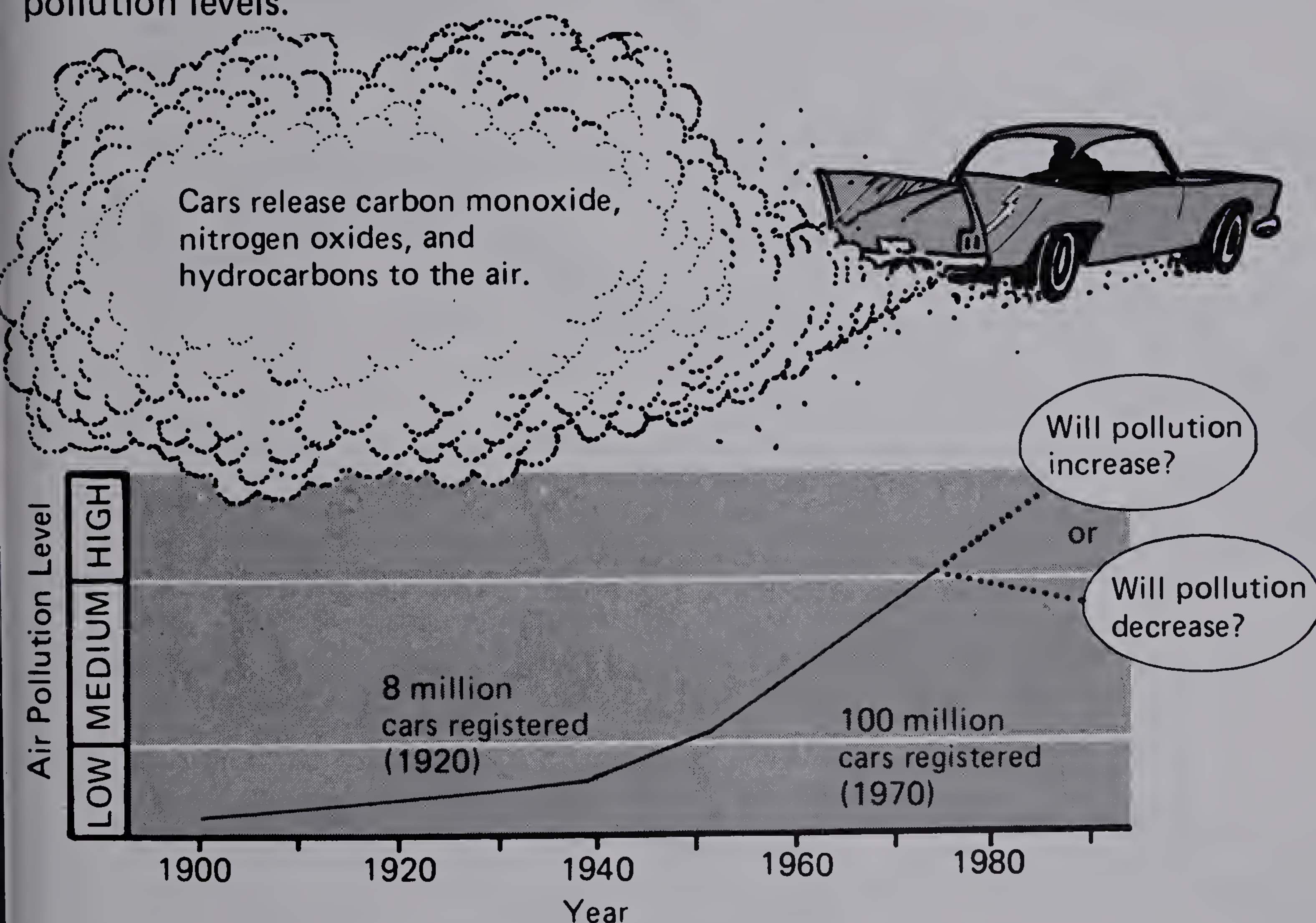


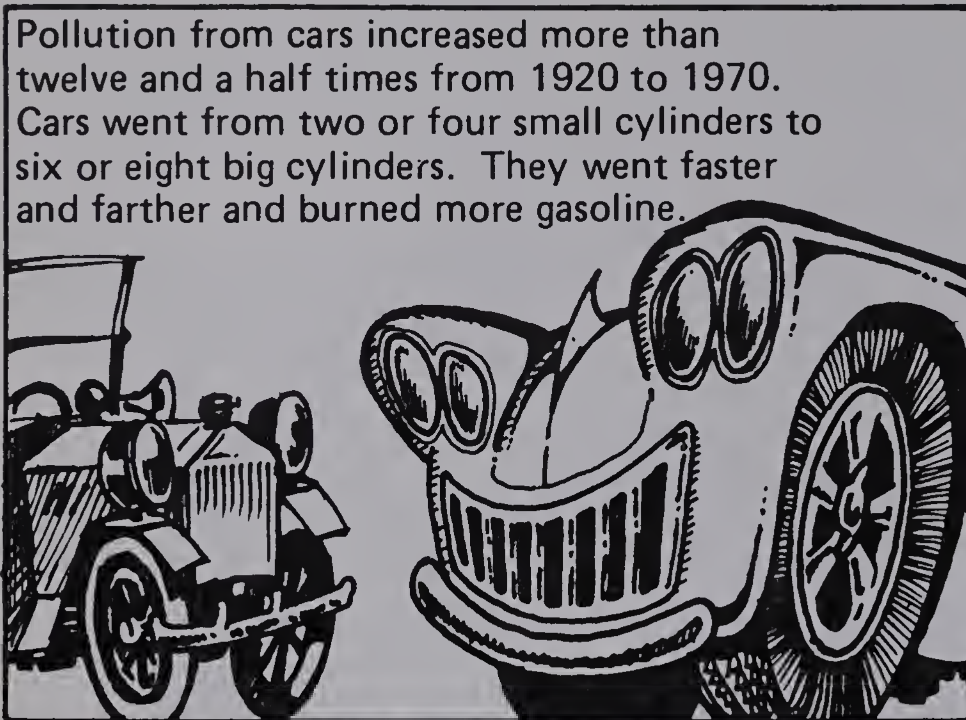
Figure 10-1

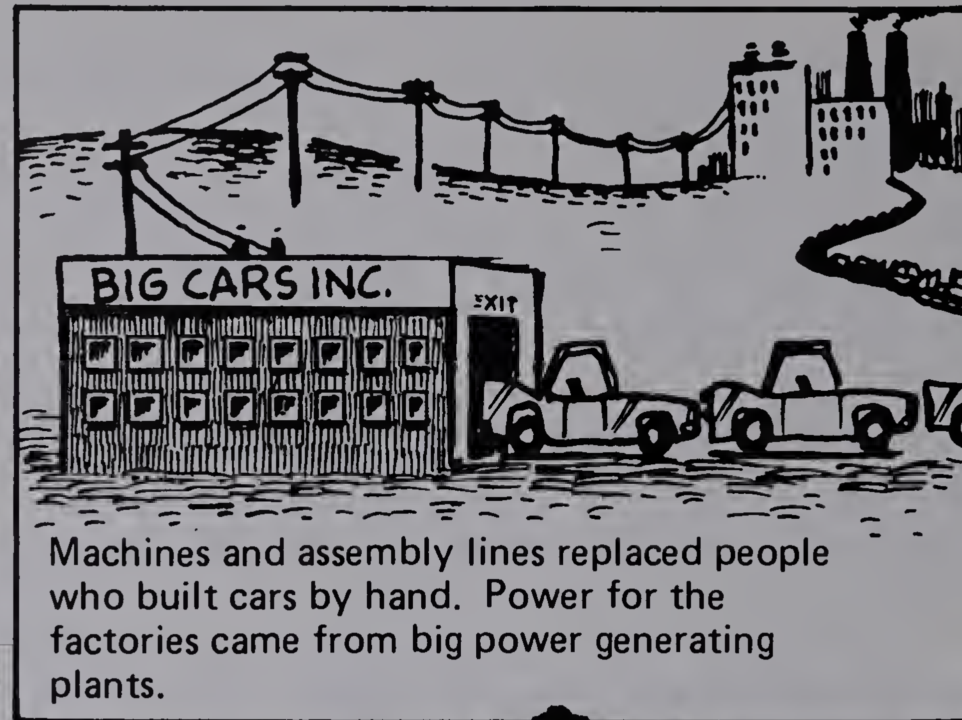
If you had trouble understanding Figure 10-1 above, read "Resource Unit 2: Reading Graphs" now.

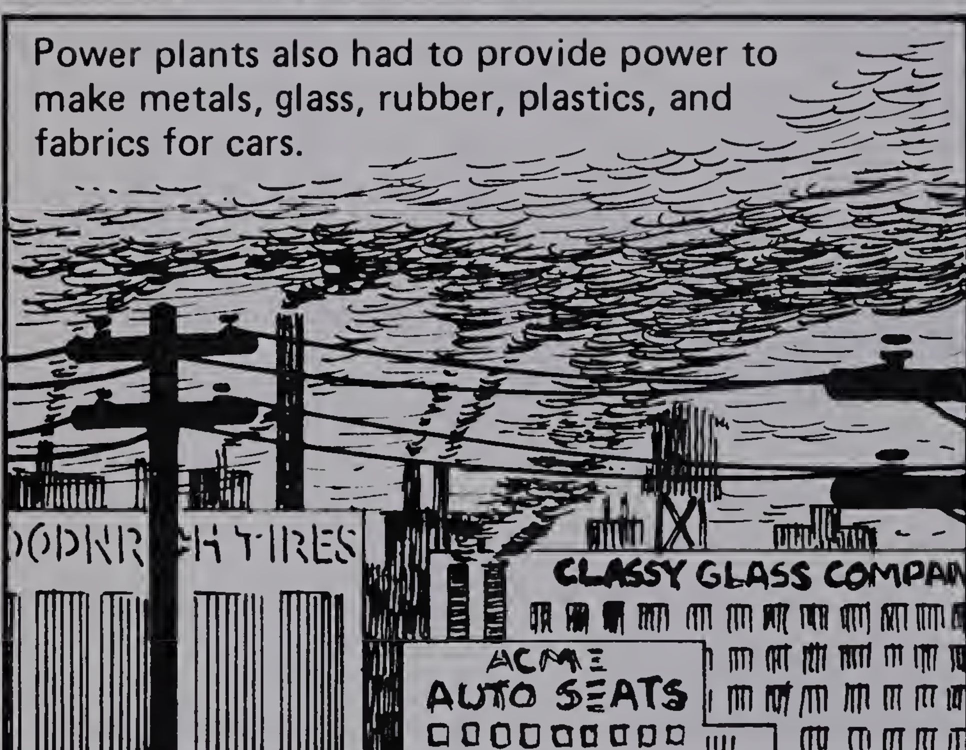
• 10-1. From 1920 to 1970, what was the trend in the amount of air pollution from cars?

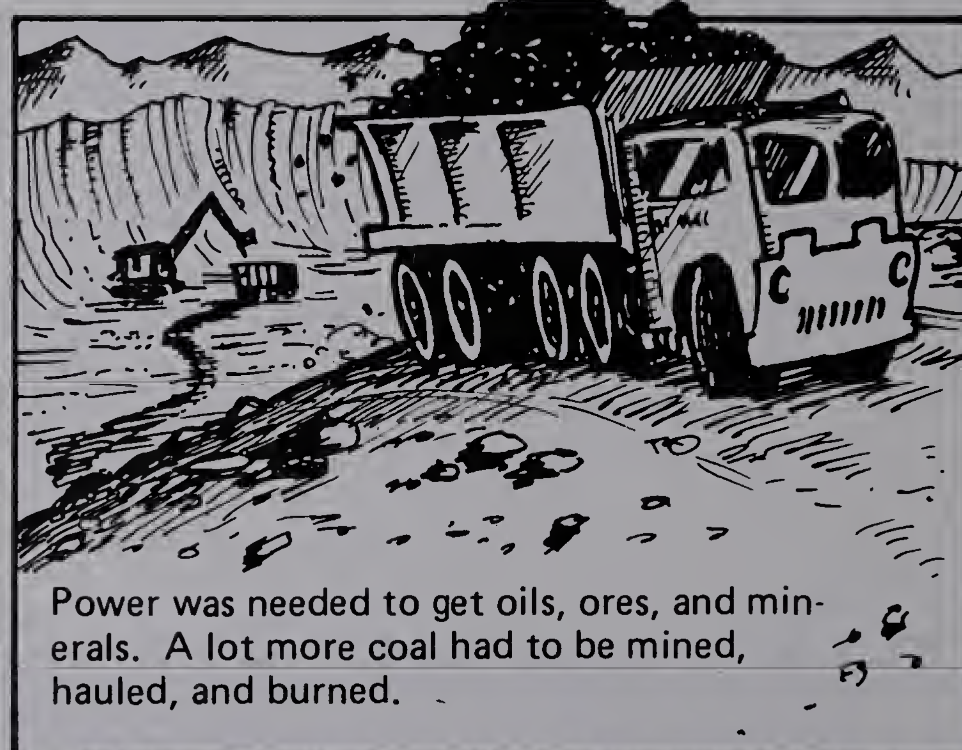
10-1. Pollution from cars was increasing.

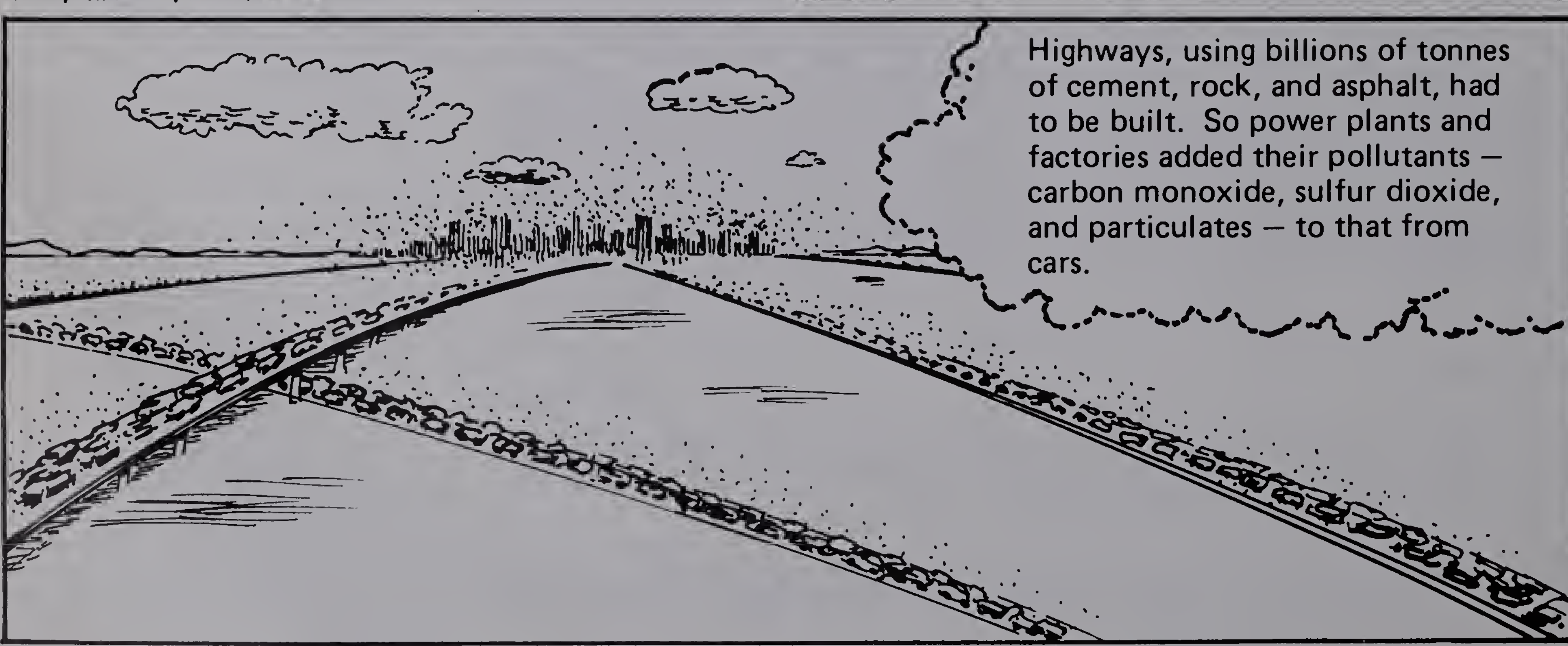
MORE CARS CAUSE INCREASES IN MANY OTHER SOURCES OF POLLUTION











• 10-2. In addition to car exhaust, what other sources of pollution increase when the number of cars increases?

10-2. Power plants, car and car-part factories, mining operations, highway construction

A lot of concern and a lot of money have been spent in recent years to try to reduce air pollution. But the problem is whether pollution will continue to increase.

• 10-3. Tell whether each of the following would cause an increase or a decrease in the pollution levels.

A. Clean-Air Acts passed by Congress to limit the amounts and kinds of pollutants that can be released to the air
B. More efficient pollution-control devices for cars and factories
C. Lower speed limits for cars
D. Smaller cars with less powerful engines

10-3. A. Decrease, B. decrease, C. decrease, D. decrease

But even if less pollution comes from each car and factory, pollution levels might still increase. Look at Figure 10-2 below.

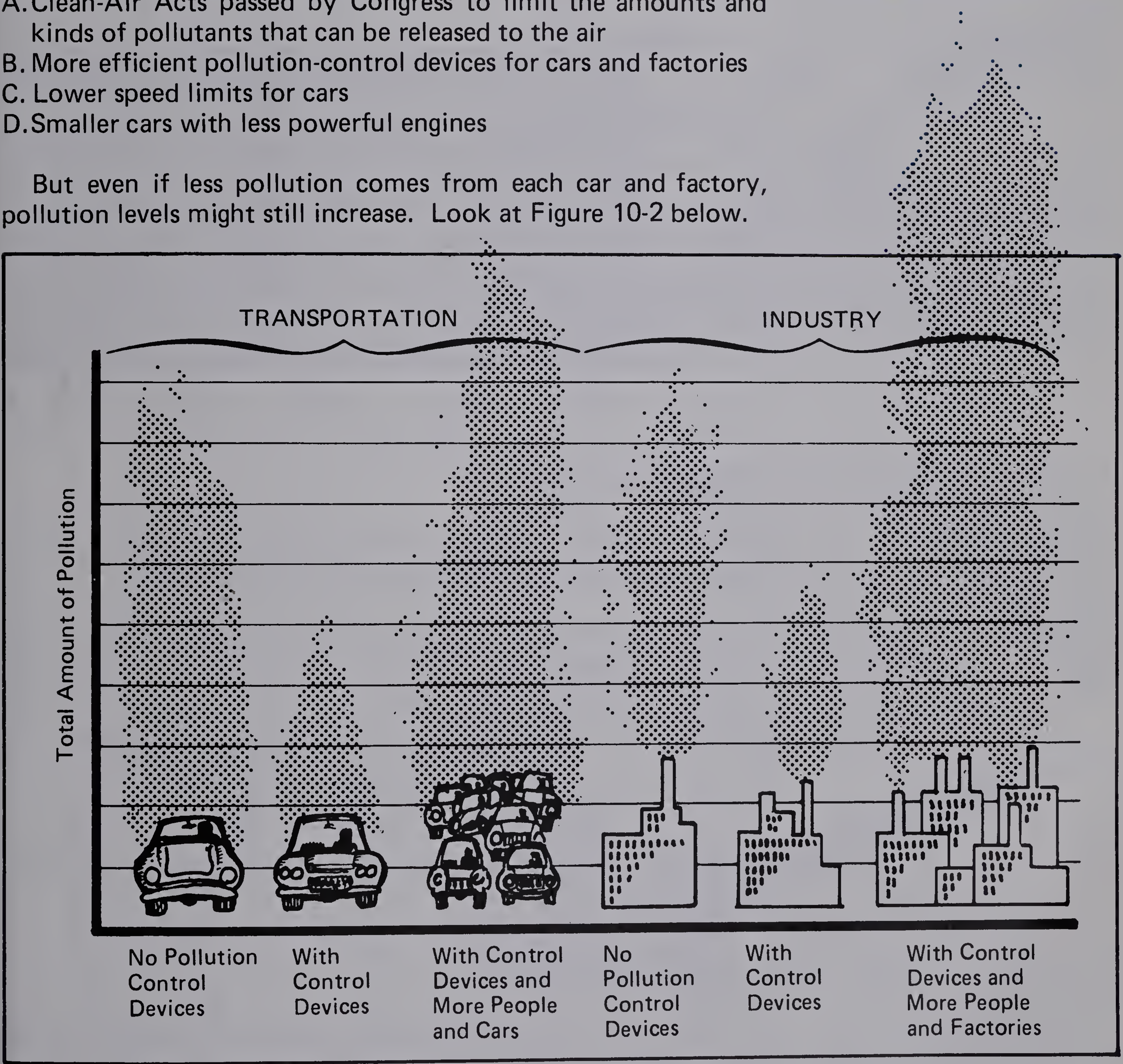


Figure 10-2

10-4. Because of more cars and more industry

● 10-4. Why might pollution levels continue to increase even with pollution-control devices?

So far, you've looked at some things that might happen to air-pollution levels and some of the reasons for those things. Now look åt some real numbers to see whether people are making any progress in controlling pollution. Figure 10-3 below shows the trend in the amount of carbon monoxide, CO, released in the United States between 1970 and 1974.

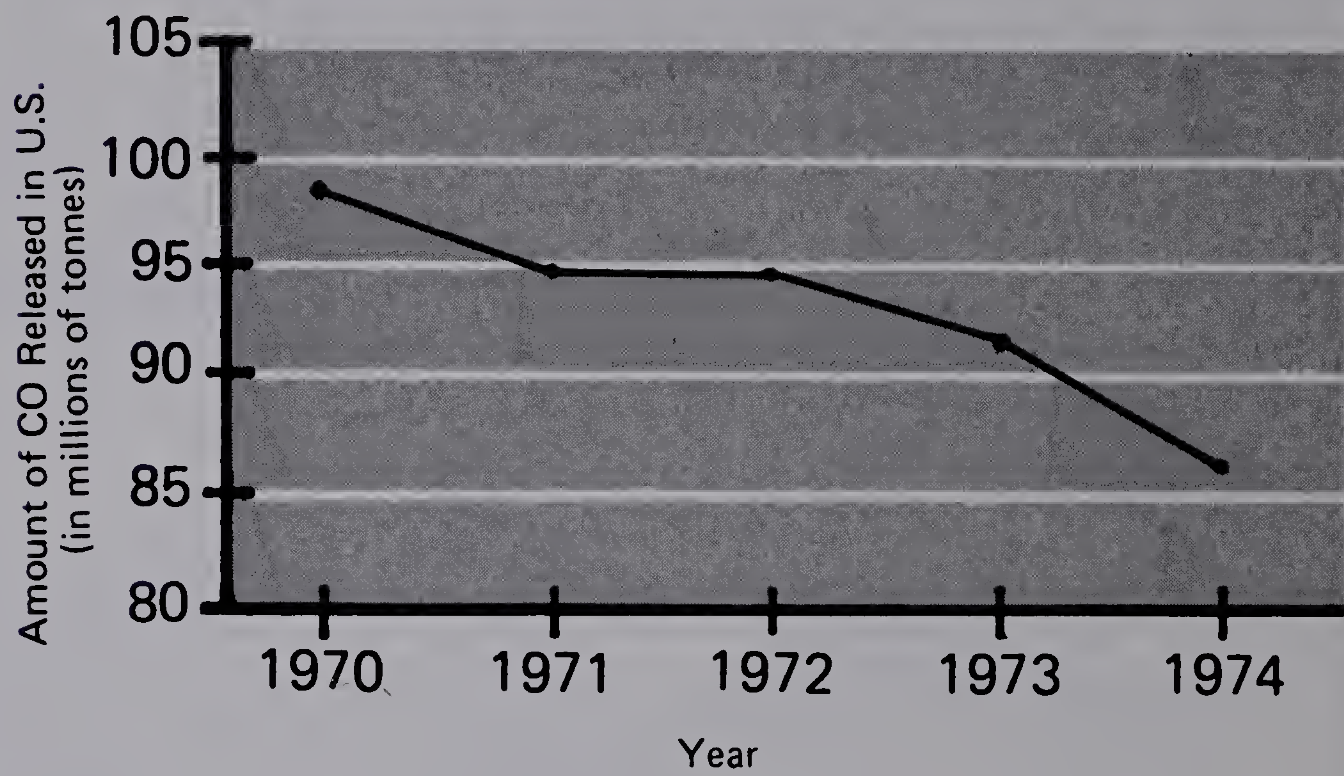


Figure 10-3

10-5. Decreasing

☆ **10-5. Between 1970 and 1974, was the amount of carbon monoxide released to the air increasing or decreasing?**

You can investigate trends in other major air pollutants. Make your own graph. You'll need the information in Figure 10-4 below and a sheet of graph paper.

	U.S. AIR POLLUTANTS RELEASED (in millions of tonnes)			
YEAR	Particulates (soot, dust)	Nitrogen Oxides	Sulfur Dioxide	Hydro-carbons
1970	25	19	31	29
1971	23	19	30	29
1972	21	20	30	28
1973	19	21	30	28
1974	18	20	29	28

Figure 10-4

If you need help making a graph, read Part A of "Resource Unit 4: Making Graphs." Then, do Step A below.

A. Draw two axes on your graph paper. Label the horizontal axis *Years*. Label the vertical axis *Pollutants (in millions of tonnes)*.

B. Mark off the horizontal scale from 1970 to 1974. Number the vertical scale from 16 to 32.

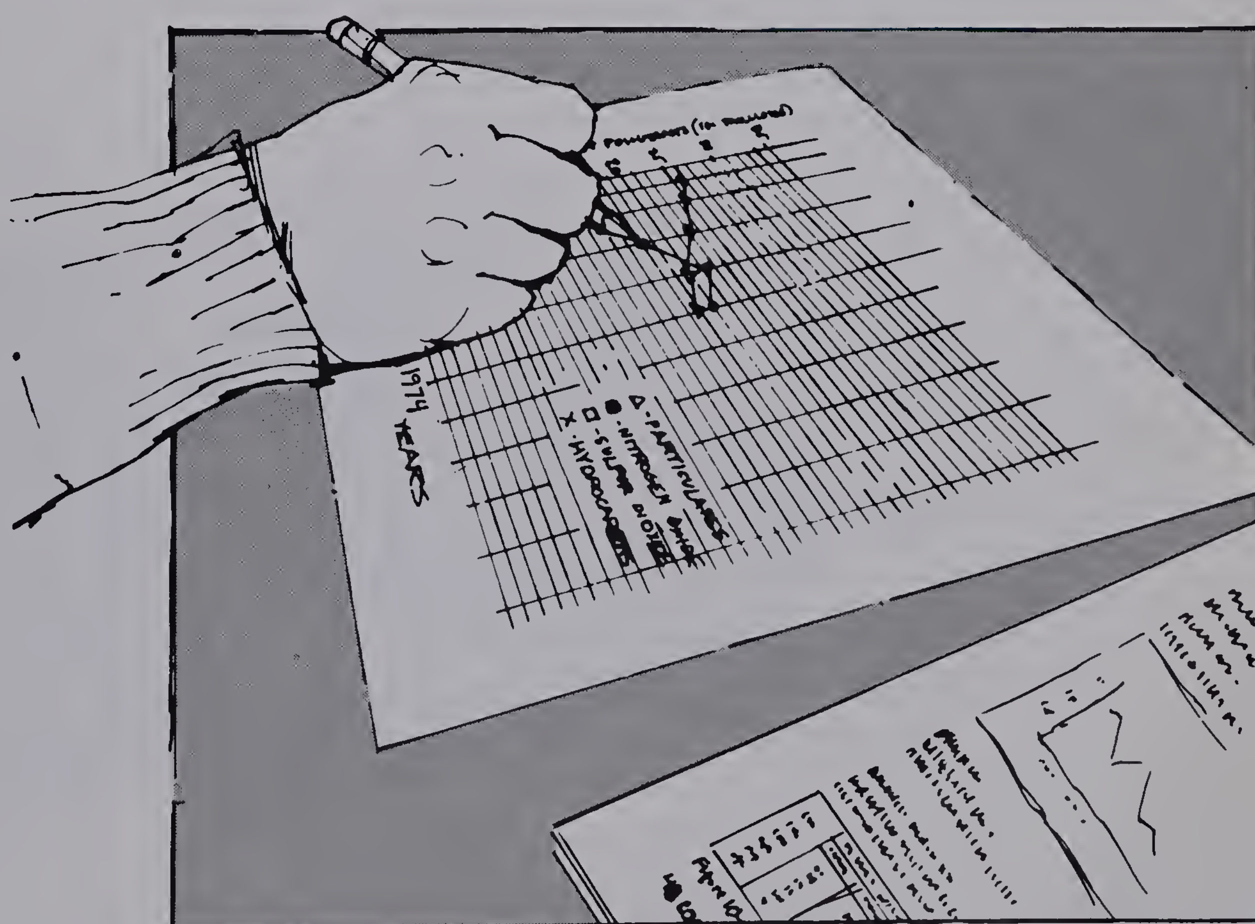
C. Plot the data for particulates from Figure 10-4 (page 48). Connect the points with straight lines as in Figure 10-3 (page 48).

• 10-6. Look at your graph. From 1970 through 1974, was the amount of air pollution from particulates increasing or decreasing?

10-6. Decreasing

D. Now, one at a time, look at the data for the rest of the pollutants in Figure 10-4 (page 48). Plot the data for each pollutant separately on your graph.

E. Label each line so that it can be clearly seen which pollutant it refers to.



• 10-7. Using your graph, describe the trend in the amount of sulfur dioxide pollution from 1970 to 1974. In nitrogen oxides. In hydrocarbons.

10-7. Decreasing; increasing; decreasing

If a decreasing trend in the release of a pollutant holds over several years, then progress has been made in controlling it. If not, progress has not been made in controlling the pollutant.

10-8. Carbon monoxide, particulates, sulfur dioxide, and hydrocarbons; nitrogen oxides

☆10-8. From your graph and the one on page 48, which of the five major air pollutants were most successfully controlled in 1974? Which were least successfully controlled?

As you have seen, since the early 1970s, people have been making some progress in the efforts to control pollution. Of the five major air pollutants, nitrogen oxides were the only ones not successfully controlled in 1974. Look at Figure 10-5 below.

YEAR	POLLUTANTS RELEASED PER HIGHWAY VEHICLE (in grams per kilometre)		
	Carbon Monoxide	Hydrocarbons	Nitrogen Oxides
1965	56	9.3	3.0
1970	49	5.0	3.3
1975	31	3.1	5.6

Figure 10-5

10-9. Carbon monoxide and hydrocarbons were decreasing. Nitrogen oxides were increasing.

☆10-9. Look at the data in Figure 10-5 above. What was the trend for each pollutant listed?

10-10. Nitrogen oxides were increasing. Carbon monoxide, sulfur dioxide, particulates, and hydrocarbons were decreasing.

☆10-10. From 1970 through 1974, which air pollutants continued to be released in increasing amounts? In decreasing amounts?

ACTIVITY 11: PLANNING

If you plan to do Activities 12 and 13, do Activity 12 first.

Activity 12 — Page 52

Objective 12-1: Explain how gases exert pressure.

Sample Question: How does a gas exert pressure on the inner wall of an automobile tire?
A. Its molecules line up and form a very heavy, solid layer against the wall.
B. Its molecules rapidly collide with the tire wall.
C. The tire wall is made of a special pressure-producing material.
D. Its molecules settle against the wall of the tire.

Objective 12-2: Explain how pressure and volume are related to each other in a gas at constant temperature.

Sample Question: If the pressure on a gas at room temperature is increased, its volume
A. increases.
B. decreases.
C. remains the same.

Activity 13 — Page 59

Objective 13-1: Describe how a gas at constant pressure changes in volume with a change in temperature.

Sample Question: For a gas held at constant pressure, a decrease in temperature will
A. increase the volume.
B. increase and then decrease the volume.
C. decrease and then increase the volume.
D. decrease the volume.

Objective 13-2: Describe how a gas held at constant volume changes in pressure with a change in temperature.

Sample Question: For a gas held at constant volume, an increase in temperature will
A. leave the pressure unchanged.
B. decrease the pressure.
C. increase the pressure.
D. increase and then decrease the pressure.

Activity 14 — Page 63

Objective 14-1: Explain the reactivities of oxygen, nitrogen, and argon in terms of their atomic and molecular structures.

Sample Question: Which statement best accounts for the way nitrogen reacts?
A. The nitrogen molecule is easily broken apart into its atoms, which can then react.
B. The triple bond in the nitrogen molecule cannot be broken, so nitrogen cannot react.
C. The nitrogen molecule is held together by a triple bond; this bond can be broken, with difficulty, and the atoms can then react.
D. The nitrogen molecule has a single bond and several unpaired electrons, allowing it to react quite readily.

Answers: 12-1. B; **12-2.** B; **13-1.** D; **13-2.** C; **14-1.** C

ACTIVITY EMPHASIS: Pressure is explained in terms of the motion of molecules. As the pressure on a gas increases, the volume decreases. When the same number of gas particles have less space to occupy, they will collide with the container walls more often. More collisions mean increased pressure.

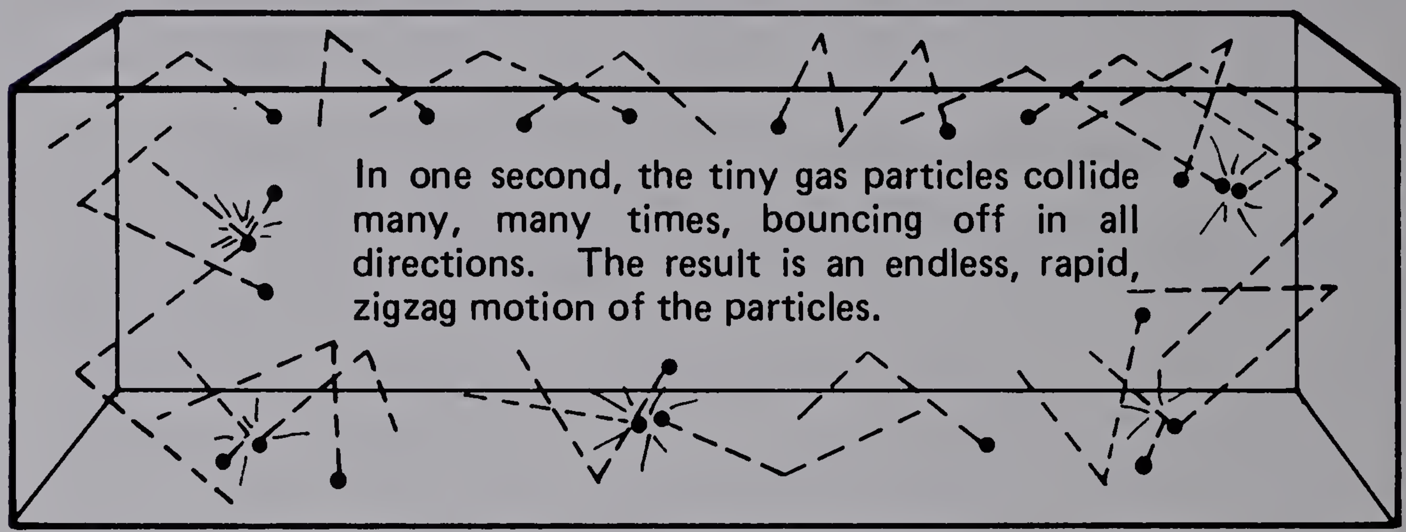
MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter. See "Advance Preparations" in ATE front matter.

If you have piston-cylinder or syringe-type gas model apparatus, it would be a good illustration for students at this point.

ACTIVITY 12: GAS PRESSURE AND VOLUME

You've seen huge trailer trucks on the road held up by nothing but the air in the tires. You know air is supporting all that weight. But it's hard to imagine how. Using a model can help.

A model will let you picture the behavior of gas particles, which are too tiny to see even under a microscope. There is evidence that these tiny gas particles have energy and are constantly moving very, very rapidly. They continually collide with one another and with the walls of any container they are in.



As each molecule of gas strikes the container wall, it exerts a push, or force — sort of like a baseball hitting a wall (or a window!). The force of each collision is very small, but there are billions and billions of tiny gas molecules in even a cubic-centimetre space, as Figure 12-1 below shows. The push of that many molecules adds up to quite a force being exerted.

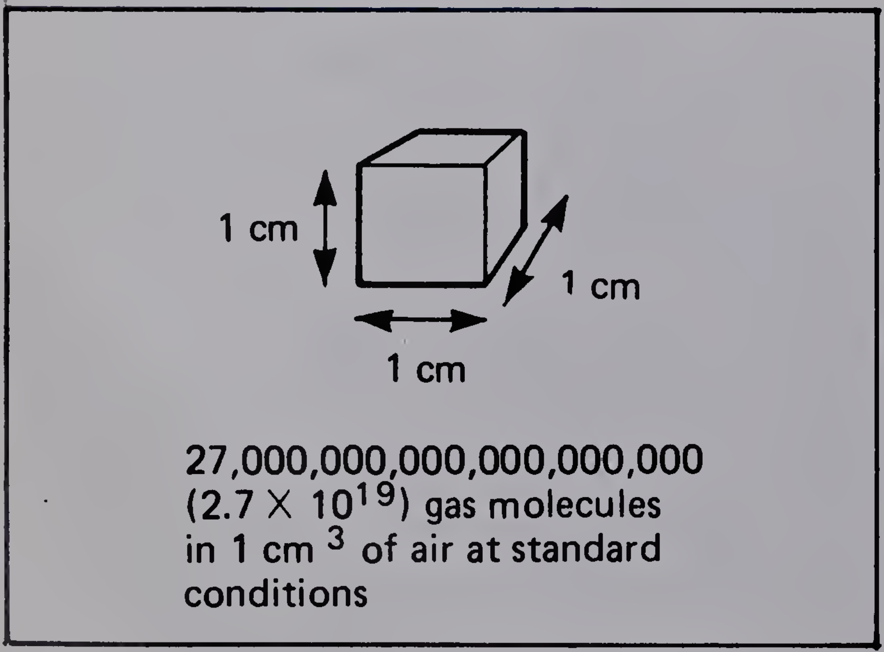


Figure 12-1

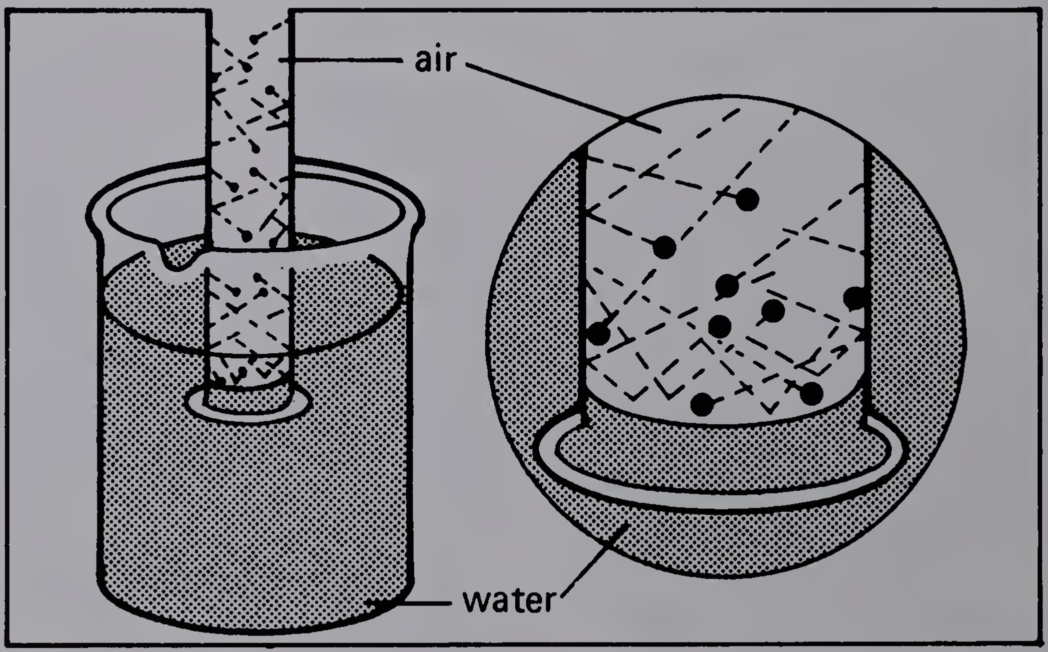


Figure 12-2

Figure 12-1 is based on Avogadro's number of molecules (6.023×10^{23}) in a mole of gas, volume = 22.4 l at 0°C and 760 mm of Hg pressure.

A test tube of air put upside down into a beaker of water can illustrate that. See Figure 12-2 above. The air molecules exert enough force in their collisions with the surface of the water to keep the water from rising in the tube. Try it if you'd like to.

• 12-1. Why doesn't the water rise in the test tube in Figure 12-2 on page 52?

12-1. The collisions of the air molecules with the surface of the water keep the water from rising in the tube.

The force with which gas molecules strike a given area is called the *pressure*. Air doesn't have to be trapped in a test tube to exert pressure. The air molecules in the atmosphere are constantly colliding with one another and with different surfaces. Thus they exert pressure on you, on your books, on the floor and ceiling and walls of your classroom — on everything.

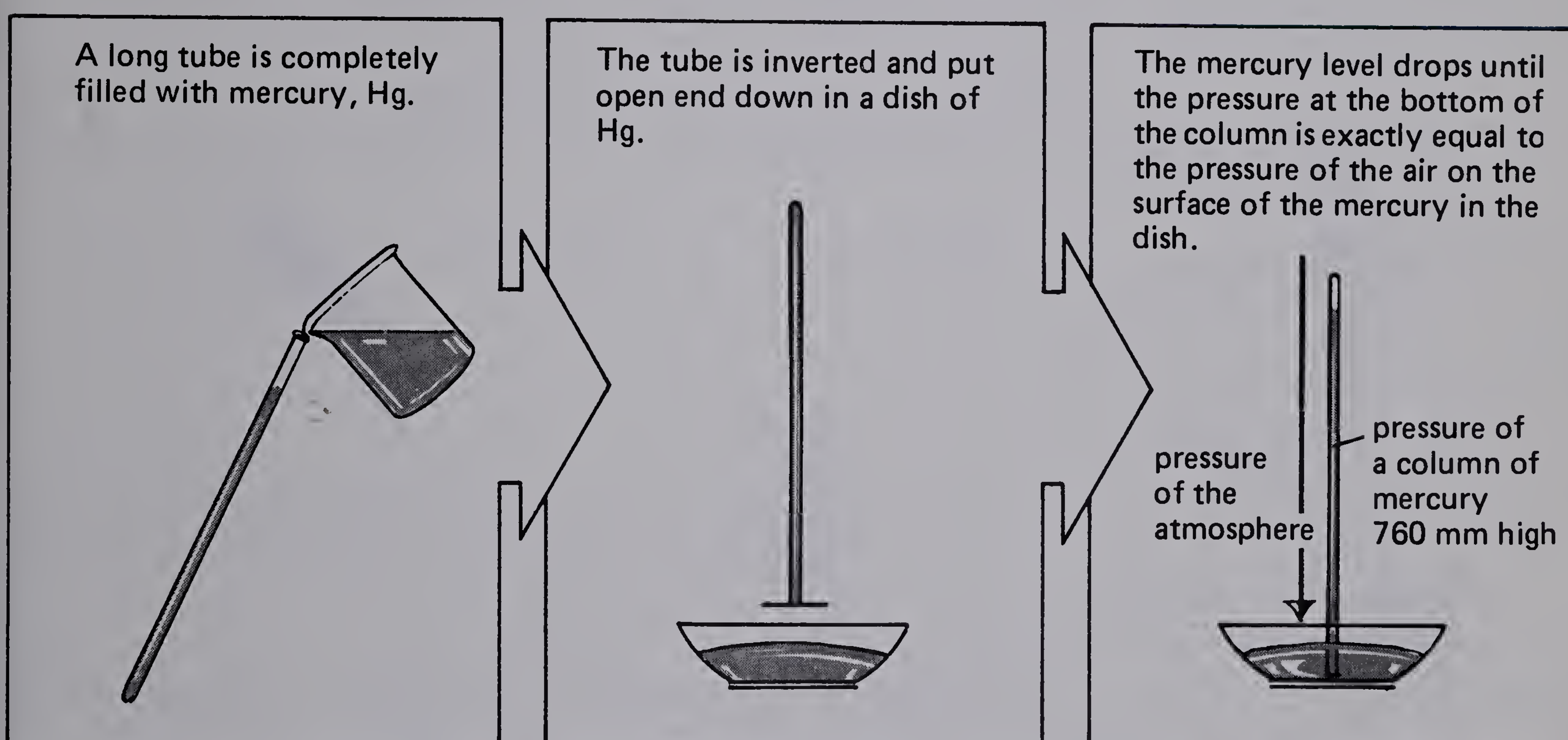
☆ **12-2. How do gases exert pressure?**

12-2. They exert pressure through the collision of their molecules with a containing wall or a surface.

• 12-3. When a truck tire is inflated, billions and billions and billions of air molecules are forced into it. How is the air in the tires able to support the truck?

12-3. The pressure resulting from the collisions of the gas molecules against the tire walls holds the walls out.

Air pressure is measured with a barometer, which can be made from a long tube filled with mercury. Figure 12-3 below shows how such a mercury barometer works.



One way to describe pressure is in *mm Hg*, that is, the height in mm of a column of Hg the gas pressure could support.

Figure 12-3

Standard atmospheric pressure is equal to 760 mm of mercury, or 760 mm Hg. Atmospheric pressure does vary, but you can get an exact measurement from a barometer.

In this activity, you'll do an investigation to find out what happens to the volume of a gas when the pressure on it changes. Before you begin, however, there are a couple of things you'll need to know. You'll be working with a sample of air trapped in the end of a glass tube. Look at Figure 12-4 below.

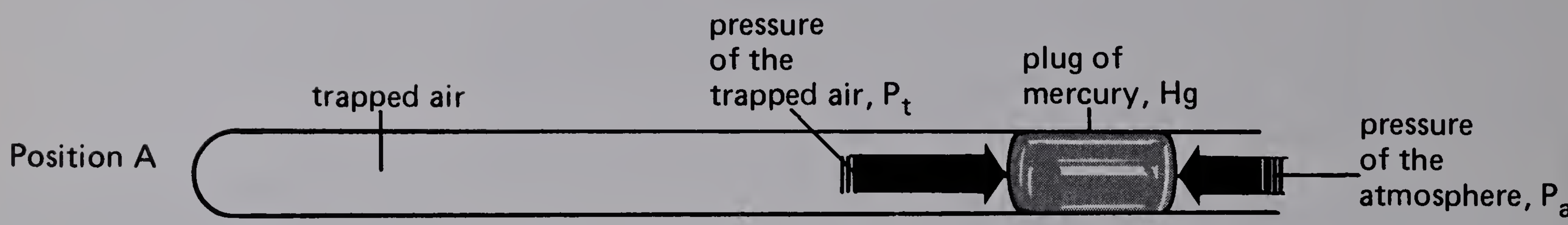


Figure 12-4

Key for Figures 12-4 and 12-5

$P_a = P_{atmosphere}$

$P_t = P_{trapped}$

$P_{Hg} = P_{Hg\ plug}$

The amount — number of molecules — of trapped air will not change. None can escape past the mercury plug. As shown in Figure 12-4 above, the trapped air is exerting pressure on the mercury plug. The trapped air molecules are constantly colliding with it. Atmospheric pressure is pushing against the other end of the Hg plug. The plug is not moving, so the push of the trapped air and the push of the atmosphere must be equal.

$$P_{trapped} = P_{atmosphere}$$

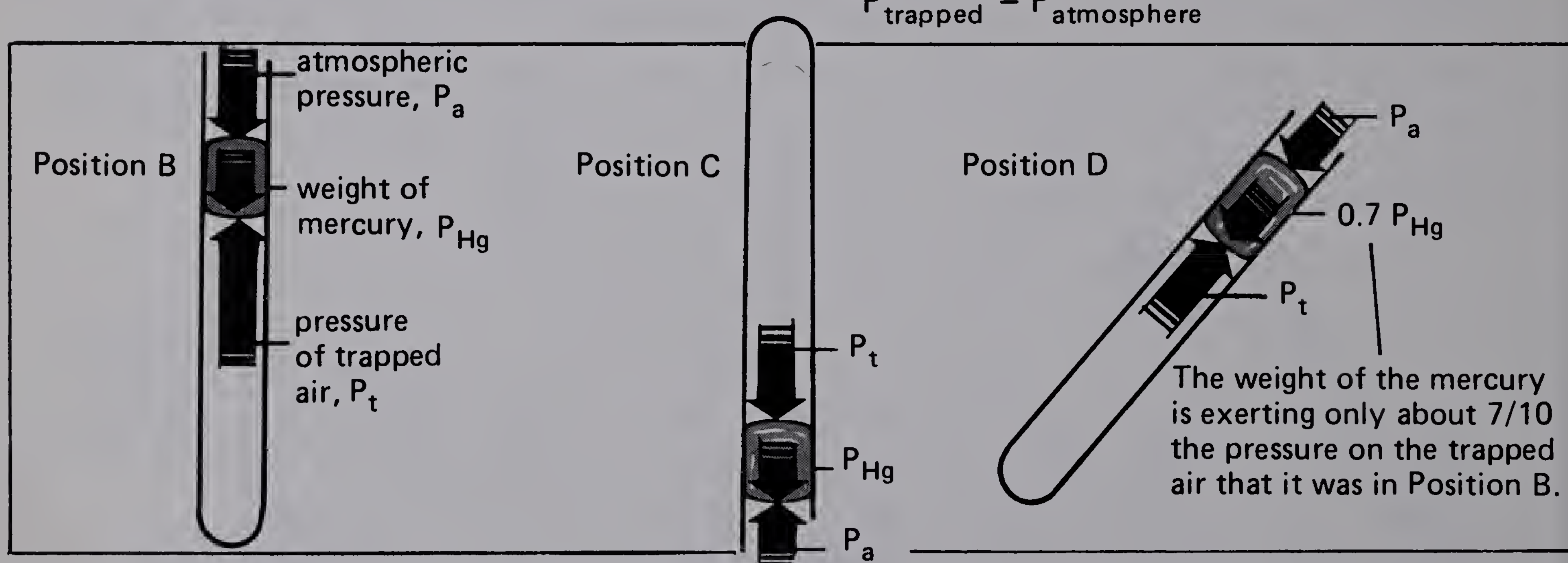


Figure 12-5

Some changes in pressure take place when the position of the tube is changed. Look at Figure 12-5 above. In Position B, the mercury plug is above the trapped air. The weight of the plug is now pushing down on the trapped air. And, as before, the atmosphere is still exerting pressure on the trapped air through the open end of the tube. Now the total pressure on the trapped air must be the atmospheric pressure *plus* the pressure of the mercury plug. And the plug is not moving, so the pressure *of* the trapped air must equal the pressure *on* the trapped air.

$$P_{trap} = P_{atmos} + P_{Hg\ plug}$$

Look at Figure 12-5 on page 54 again. In Position C, the tube has been turned upside down.

● 12-4. In Position C, is the Hg plug pushing against the trapped air or pushing with it against the pressure of the atmosphere?

12-4. The plug is pushing with the trapped air against the pressure of the atmosphere.

Now, the pressure of the trapped air, along with the pressure of the mercury plug, is equal to the pressure of the atmosphere.

$P_{trap} + P_{Hg\ plug} = P_{atmos}$ OR $P_{trap} = P_{atmos} - P_{Hg\ plug}$

In Position D in Figure 12-5 on page 54, notice that the weight of the mercury plug is only partially pushing on the trapped air.

● 12-5. What would be the pressure on the trapped air in Position D?

12-5. $P_{trap} = P_{atmos} + 0.7\ P_{Hg\ plug}$

In your investigation, you'll change the position of the tube with trapped air and determine what happens to the volume of this air as the pressure is changed. You will use the *length* of the trapped air column as an indirect measure of its volume. The length of the mercury plug in millimetres (mm) is a direct measure of the pressure the plug exerts. You'll need the following.

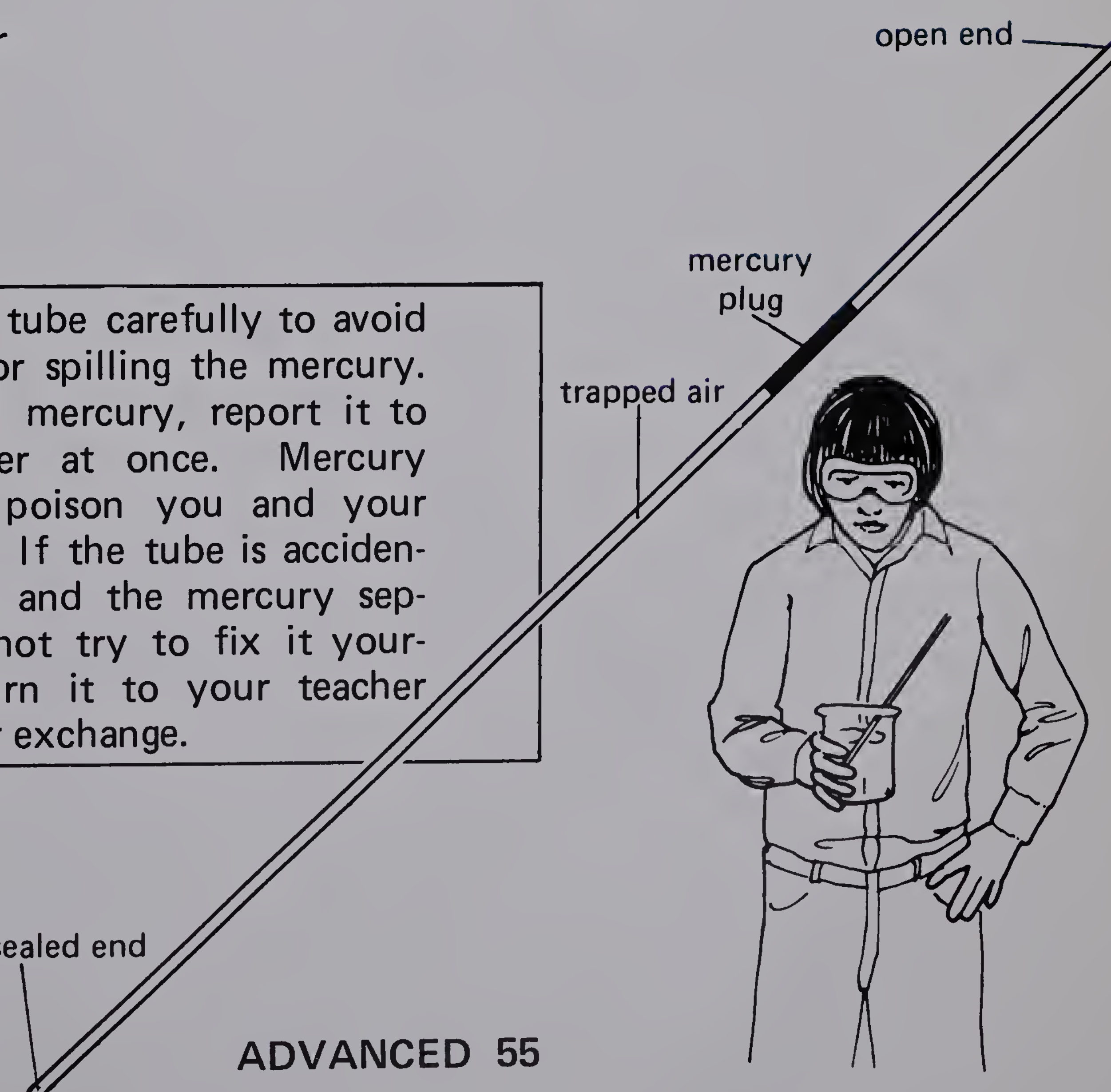
See the special procedures for handling mercury spills in "Background Information" in the ATE front matter.

A plot of volume versus pressure will verify Boyle's law: $PV = k$.

safety goggles
special gas tube in 600-ml beaker
metric ruler
classroom barometer
triangle, 45°, 45°, 90°
graph paper

CAUTION	Handle the tube carefully to avoid separating or spilling the mercury. If you spill mercury, report it to your teacher at once. Mercury vapor can poison you and your classmates. If the tube is accidentally jarred and the mercury separates, do not try to fix it yourself. Return it to your teacher for repair or exchange.

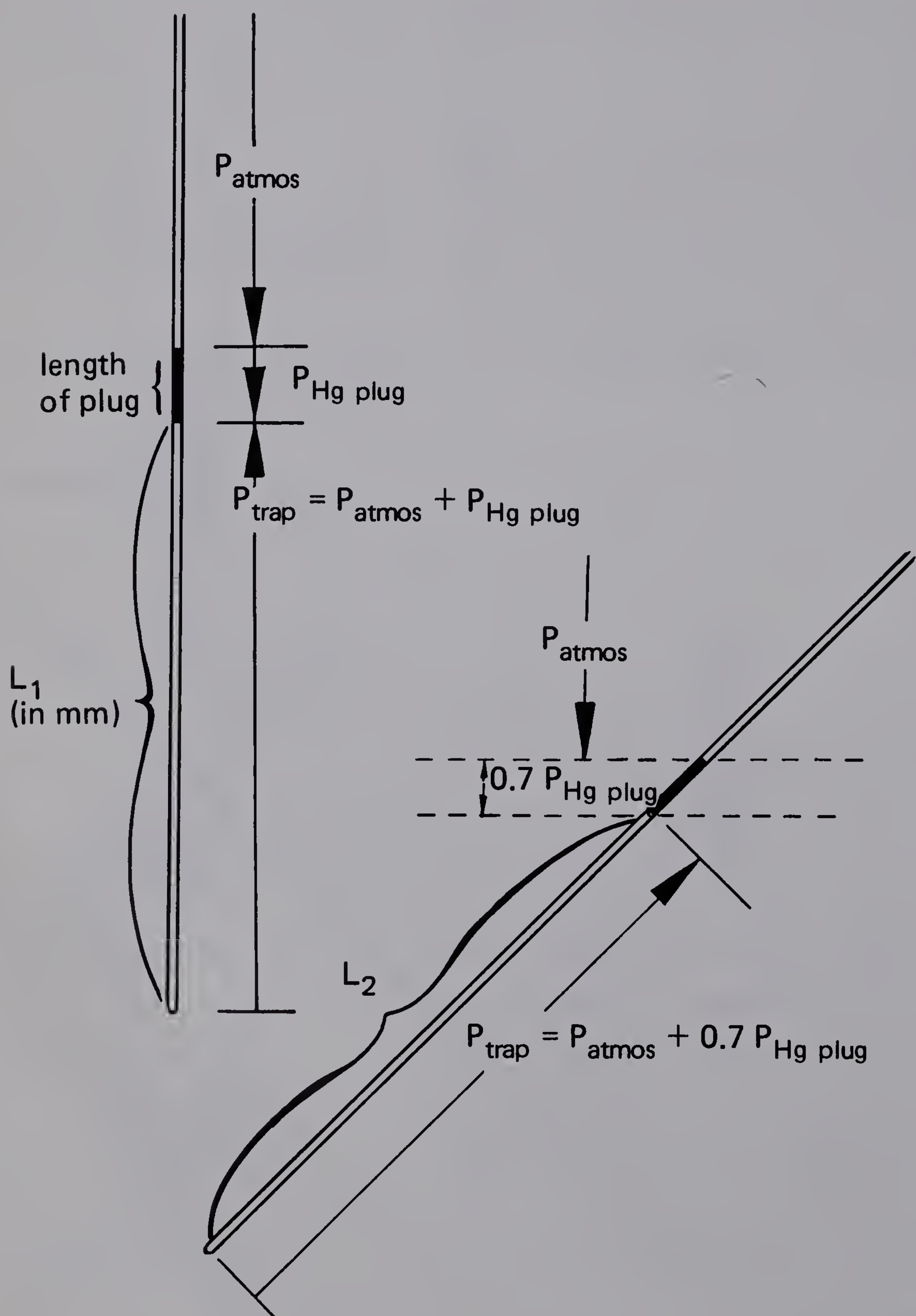
A. Carefully examine the tube in the beaker. It should contain a column of air trapped beneath a mercury plug.



P_{atmos} = ______ mm Hg (from barometer) $P_{Hg\ plug}$ = ______ mm Hg (length of mercury plug)		
POSITION OF TUBE	LENGTH OF TRAPPED AIR (mm)	TOTAL PRESSURE ON TRAPPED AIR (mm Hg)
Step E. Open end up	L_1 = ______	$P_{atmos} + P_{Hg\ plug}$ = ___ + ___ = ___
Step F. Tilted 45° from vertical	L_2 = ______	$P_{atmos} + 0.7\ P_{Hg\ plug}$ = ___ + ___ = ___
Step G. Horizontal	L_3 = ______	P_{atmos} = ___
Step H. Tilted 45° down	L_4 = ______	$P_{atmos} - 0.7\ P_{Hg\ plug}$ = ___ − ___ = ___
Step I. Closed end up	L_5 = ______	$P_{atmos} - P_{Hg\ plug}$ = ___ − ___ = ___

B. Copy this table into your notebook.

C. From the classroom barometer, read the atmospheric pressure, P_{atmos}. Record this pressure in millimetres of mercury (mm Hg) at the top of your table.

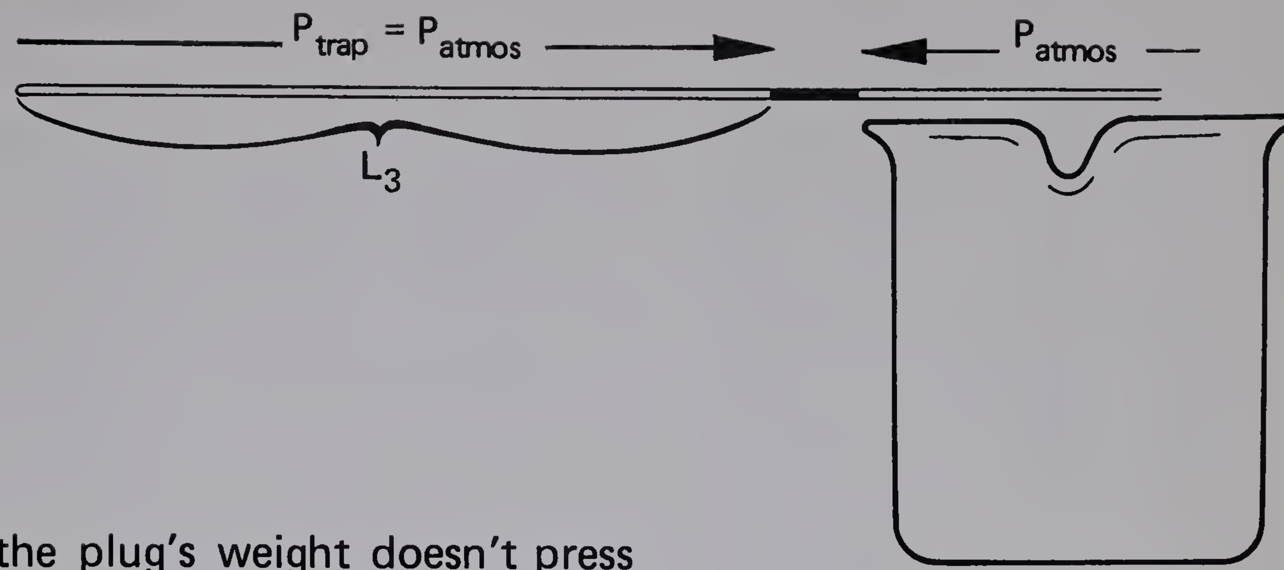


D. Carefully hold the gas tube vertically over the beaker with the open end up. Measure the length of the mercury plug in millimetres. This is a direct measure of the pressure exerted by the plug, $P_{Hg\ plug}$. Record this in the top of your table.

E. Measure the length of the trapped air column, L_1, in millimetres. Record the results in your table. Calculate in your table the pressure of the trapped air, P_{trap}.

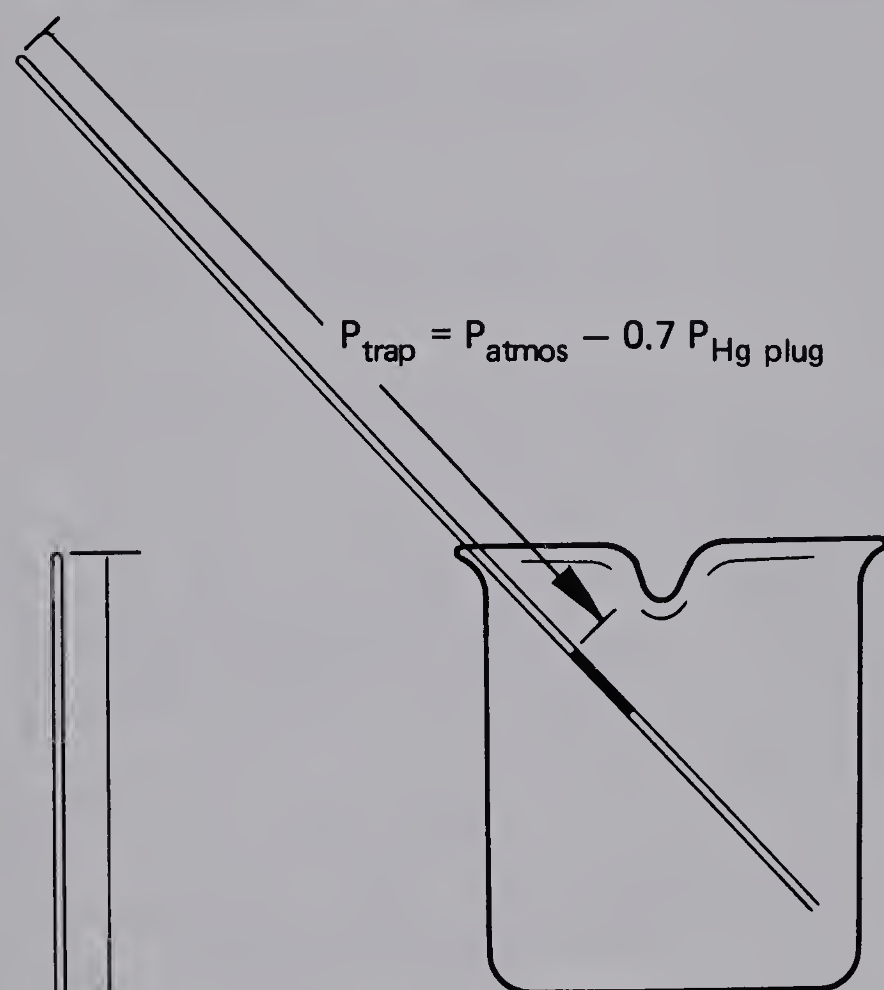
F. Carefully and slowly, tilt the gas tube until it is 45 degrees from the vertical. Use the triangle to judge the 45° angle. Measure the length of the trapped air column, L_2, in millimetres. Record it in the table. Calculate P_{trap} in this position.

G. Tilt the gas tube gently and slowly until it is horizontal and the open end is over the beaker. Measure the length of the trapped air column, L_3. Record it in the table.

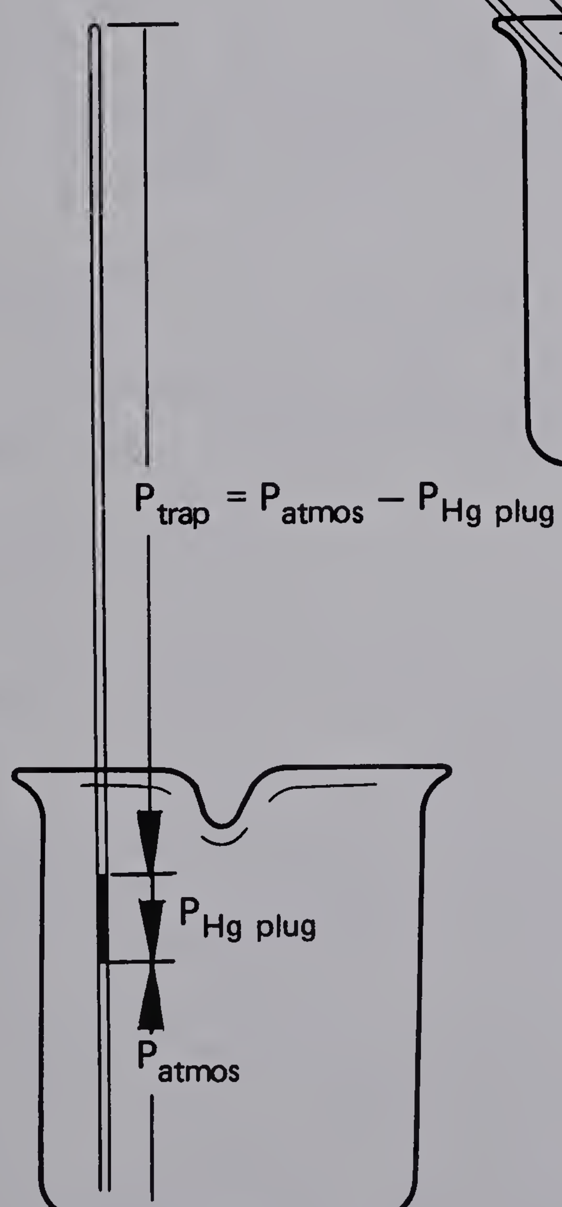


When the tube is horizontal, the plug's weight doesn't press on the trapped air.

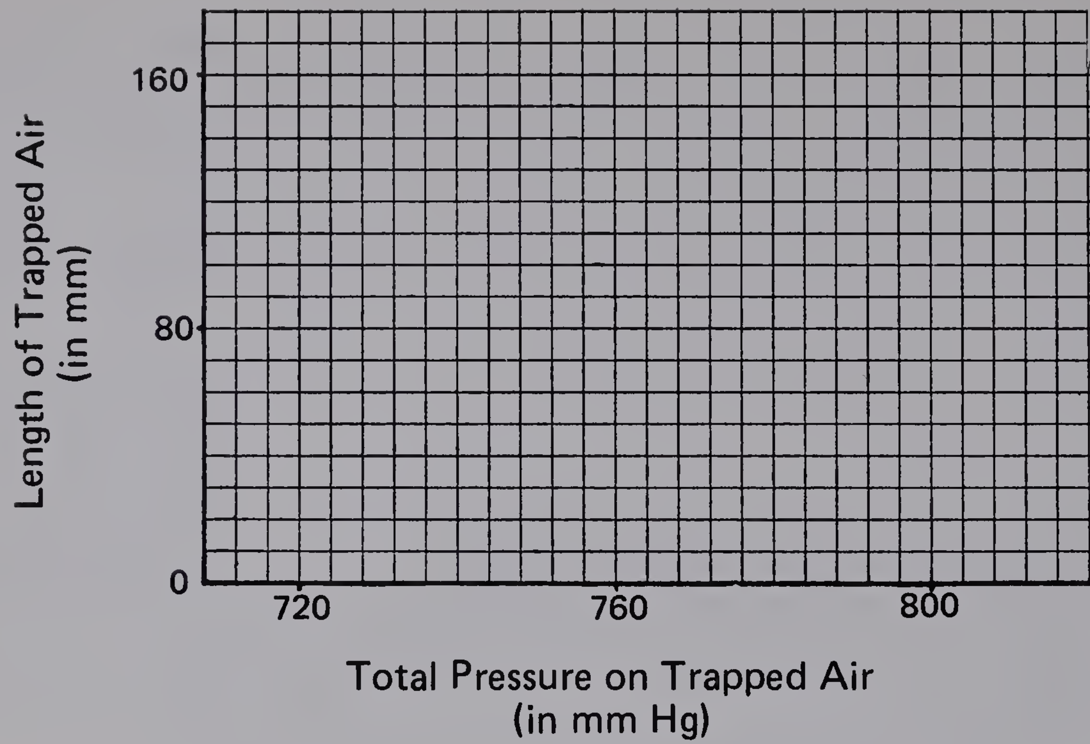
H. With the open end of the gas tube over the beaker, gently and slowly tilt the tube to a 45 degree angle pointing down. Use the triangle to judge the angle. Without jarring it, measure the length of the trapped air column, L_4. Record L_4 in your table. Calculate P_{trap} in this position.



I. Slowly and carefully, without jarring it, tilt the gas tube to the vertical with the closed end up. Measure and record the length of the trapped air column, L_5. Calculate P_{trap} in this position.



J. Very carefully, rotate the tube out of the beaker until the open end is up. Keeping the open end up, put the gas tube into the beaker in the position shown in Step A.



K. Using graph paper, make a graph like the one shown. Graph the results from your table by plotting each pair of length—pressure (of trapped air) points. Then connect the points.

12-6. The volume decreased.

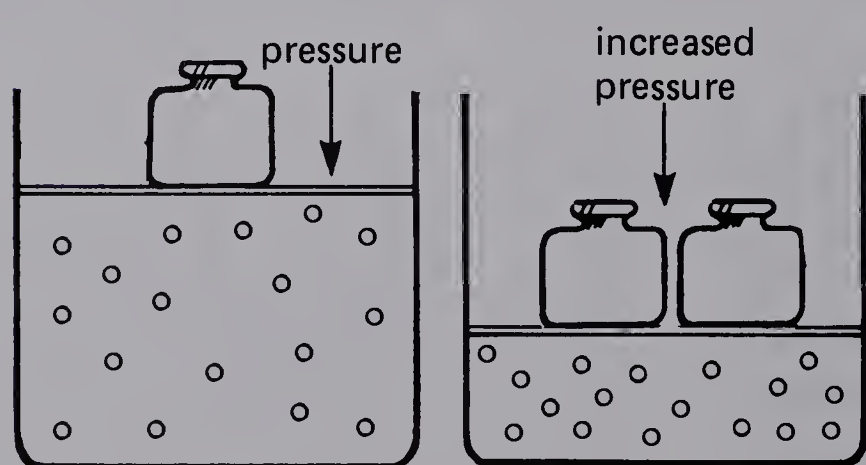
12-7. The pressure increased.

[Graphed results may vary. (There's usually a bigger problem with the 45° points than with the vertical or horizontal points. But a protractor isn't needed for the accuracy desired. Usually, results approximate PV = k.)]

● 12-6. Use your graph to determine what happened to the volume of the trapped gas as the pressure on it was increased.

● 12-7. Now use your graph to determine what happened to the pressure on the trapped gas as its volume decreased.

The number of molecules of gas trapped in the tube did not change. But the pressure of the gas did, and so did the volume occupied by the gas molecules. The model in Figure 12-6 below helps explain these changes.



The volume is decreased. The molecules are pushed closer together. They will have more collisions.

Figure 12-6

12-8. The number of collisions will increase.

☆ **12-8. If the pressure on a gas is increased, the molecules are pushed closer together. What effect will this have on the number of collisions the molecules will have?**

12-9. The volume increases.

☆ **12-9. As the pressure on a trapped gas is decreased, what happens to its volume?**

12-10. The pressure must be increasing.

☆ **12-10. When the volume of a gas decreases, what must be happening to its pressure?**

ACTIVITY 13: HOT AIR

ACTIVITY EMPHASIS: The effect of temperature on gas volume is measured at constant pressure. The effect of temperature on pressure at constant volume is discussed.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter. See "Advance Preparations" in ATE front matter.

Activity 12 discusses a gas model that says gases are made of tiny particles constantly bouncing around. With that model, you can make some predictions about some other behaviors of gases.

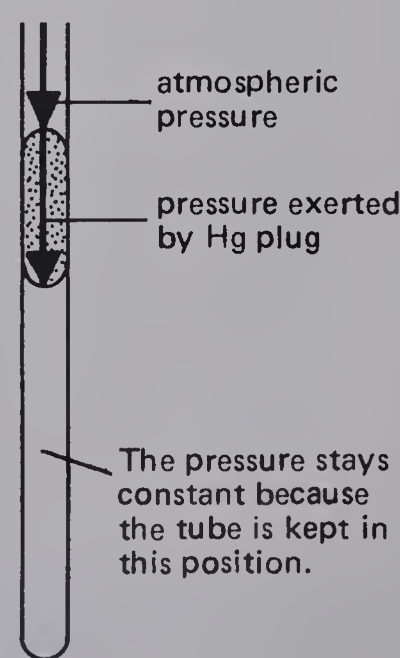
Suppose the gas molecules are given more energy and so move faster. How would that affect their volume? You can make gas molecules move faster by heating them. To see what effect that has on volume, you will need these materials.

- safety goggles
- Bunsen burner
- thermometer and gas tube assembly
- 30 cm of wire
- 10 cm of wire
- graph paper
- metric ruler
- large test tube
- ring stand
- 2 test-tube clamps

See the special procedures for handling mercury spills in "Background Information" in the ATE front matter.

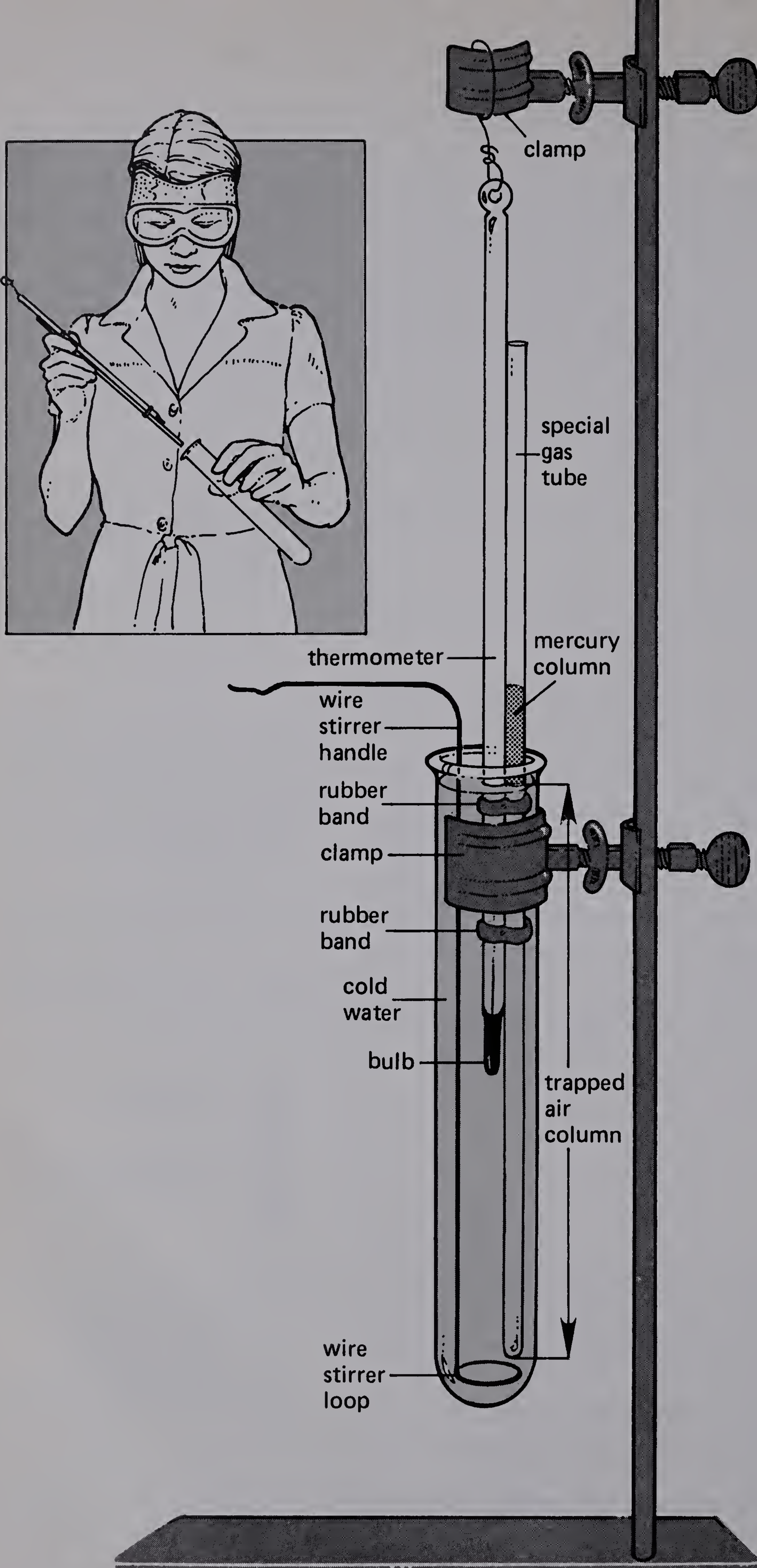
A. In your notebook, make a table like this. Leave about seven or eight lines for entries.

TEMPERATURE (in °C)	LENGTH OF AIR COLUMN (in mm)



As long as the assembly that you'll work with is kept vertical during heating, the trapped gas will be under constant pressure.

CAUTION	If the mercury is separated or if you accidentally jar the tube and separate the mercury, do not try to fix it yourself. Return the tube to your teacher for repair or exchange. If you spill any mercury, tell your teacher immediately.

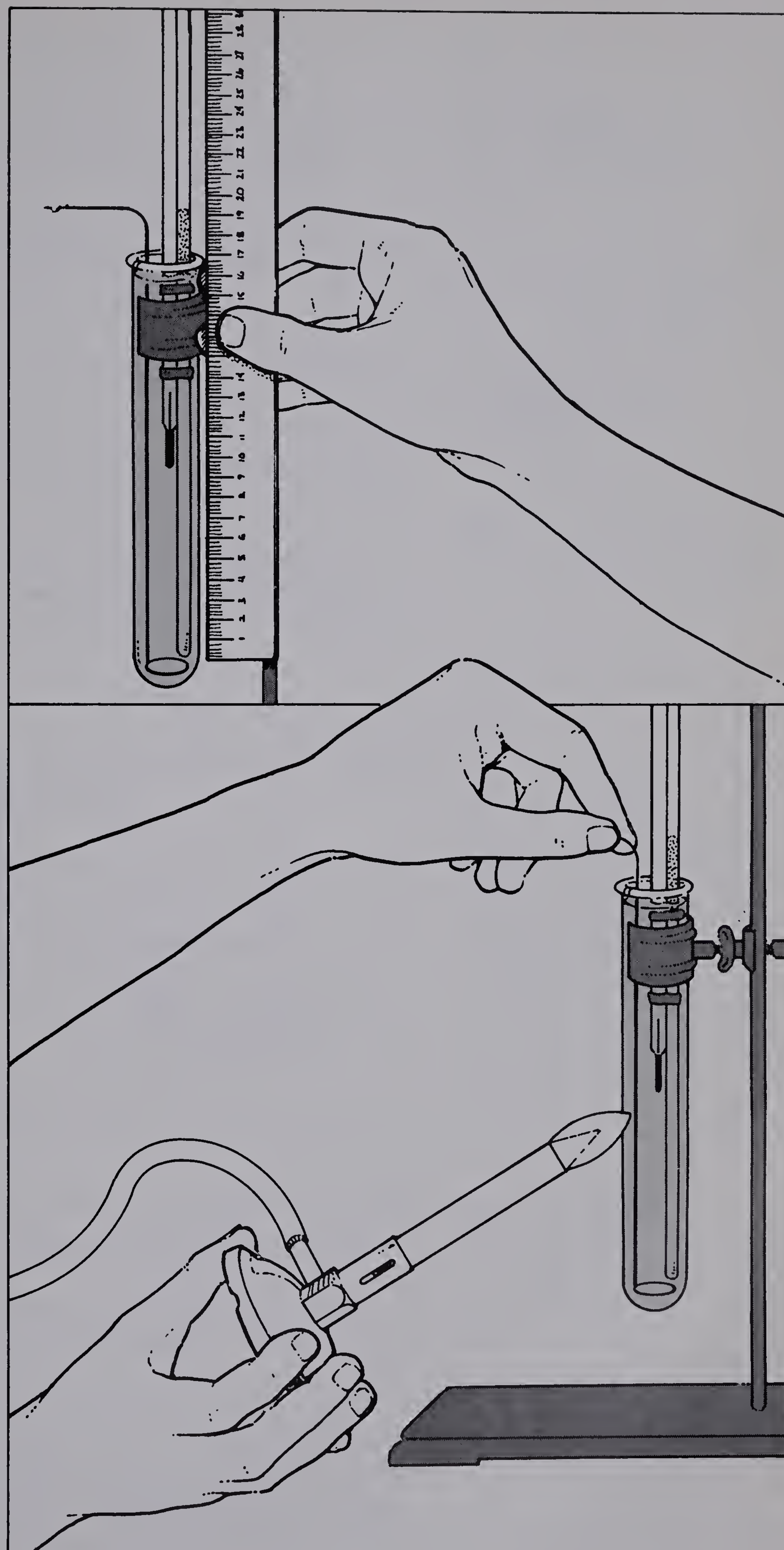


B. Attach the 10-cm piece of wire to the thermometer. Suspend the thermometer and gas-tube assembly in the test tube. Clamp the test tube to the ring stand.

C. Bend the 30-cm piece of wire to form a stirrer with a loop and handle, as shown. Move the wire loop up and down to be sure it does not rub the test tube or the thermometer and gas-tube assembly. The wire loop is for stirring the water to assure an even temperature throughout.

D. Fill the test tube to within 2 or 3 cm of the top with cold water. Stir the water until the temperature remains constant.

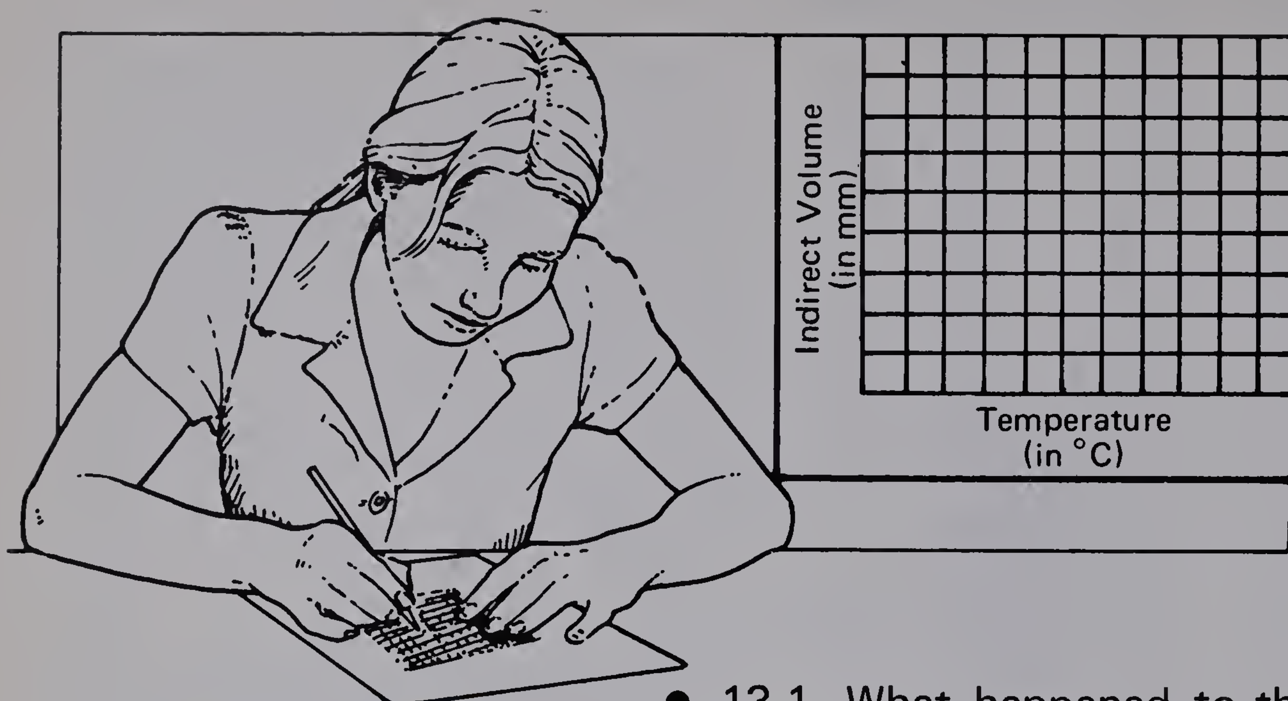
E. Measure and record both the temperature and the length of the trapped air column.



F. Very gently warm the test tube by brushing a burner flame up and down its side. Stir the water slowly. When the temperature has increased about 5°C, stop heating the test tube. Continue stirring the water until the temperature stops increasing. Then, very quickly measure and record the length of the trapped air column and the temperature. The temperature will fall rapidly.

G. Continue to heat and stir the water. Measure and record the length of the air column and the temperature at about five-degree intervals until the temperature reaches about 95°C.

H. Graph the data from your table. Label the horizontal axis *Temperature (in °C)* and the vertical axis *Indirect Volume (in mm)* — the length of the air column.



13-1. The volume increased.

With temperature in degrees Kelvin (°C + 273), the plot verifies Charles's Law: $\frac{V}{T} = k$.

● 13-1. What happened to the volume of the trapped gas as its temperature was raised?

At constant pressure, you put more energy (heat) into the molecules of the trapped air. The volume changed. The model in Figure 13-1 below helps explain what was happening.

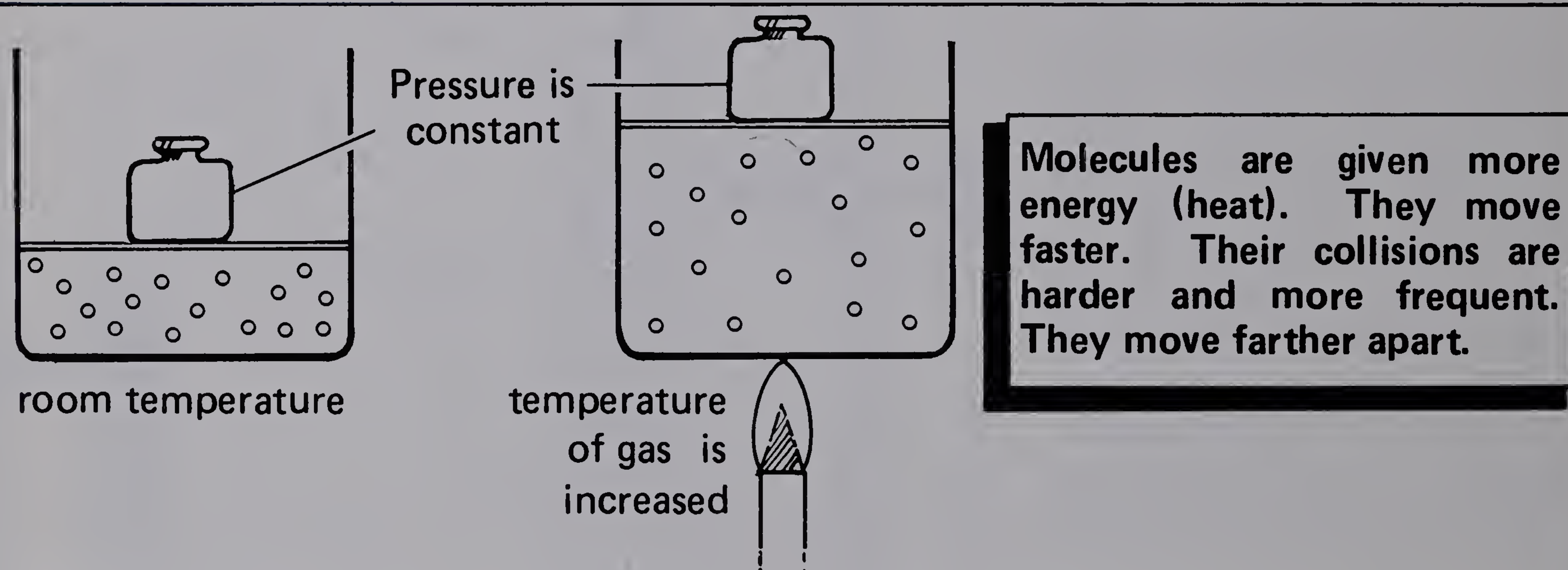


Figure 13-1

13-2. The molecules are given more energy by heating, so they move faster and farther apart. Hence, the volume increases.

13-3. The temperature must be decreasing.

☆ **13-2. Explain, in terms of molecules, what happens to the volume of a gas as it is heated.**

☆ **13-3. At constant pressure, if the volume of a gas decreases, what must be happening to its temperature?**

Heating can't change the volume of a gas in a rigid container. But the speed of the molecules is still increased by the heat. And the molecules bounce harder and more often against the walls of the container. So the pressure inside the container goes up. If it goes up too much, it can tear the container apart.

On a long trip at high speeds, car tires get heated because of friction.

• 13-4. Suppose tires were made of a material that didn't stretch. What would happen to the air pressure inside them during a long trip?

13-4. During a long trip, the air pressure would increase.

• 13-5. Since tires can stretch somewhat, what happens to their size when the air inside them is heated?

13-5. Their size (the volume of the air inside) increases.

What actually happens in tires is a combination of changes in volume and pressure. Tires are somewhat flexible, so there is some increase in the volume of the air inside. But because tires are not completely flexible, the increased speed of the air molecules also causes some increase in pressure. A properly inflated tire prevents excessive pressure and heat buildup. If a tire has too much air pressure to start with, a little heat makes the pressure too high. If a tire has too little air in it, friction overheats and weakens the rubber.

ACTIVITY 14: REACTIVITY EXPLAINED

ACTIVITY EMPHASIS: The differences in chemical reactivity of argon, nitrogen, and oxygen can be explained by differences in electron structure and in the bonds in diatomic molecules of N_2 and O_2.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter.

Substances don't all behave the same way chemically, and the gases in air are no exception. Some are more reactive than others. The three most abundant gases in air — nitrogen, oxygen, and argon — behave very differently from one another.

You can see the difference in behavior between oxygen and nitrogen directly. Since air is mostly nitrogen and oxygen, you can heat magnesium metal with air and see which gas reacts. You can tell quite easily which gas reacts more readily.

When oxygen reacts with magnesium metal, a white solid compound is formed.

$2\,Mg$	+	O_2	$\longrightarrow$	$2\,MgO$
magnesium (metal)		oxygen (gas)		magnesium oxide (white solid)

When nitrogen reacts with magnesium metal, a green solid compound is formed.

$3\,Mg$	+	N_2	$\longrightarrow$	Mg_3N_2
magnesium (metal)		nitrogen (gas)		magnesium nitride (green solid)

Since the products of the two reactions are different colors, you can tell which gas reacted by the color of the product formed. Try it. You will need the following materials.

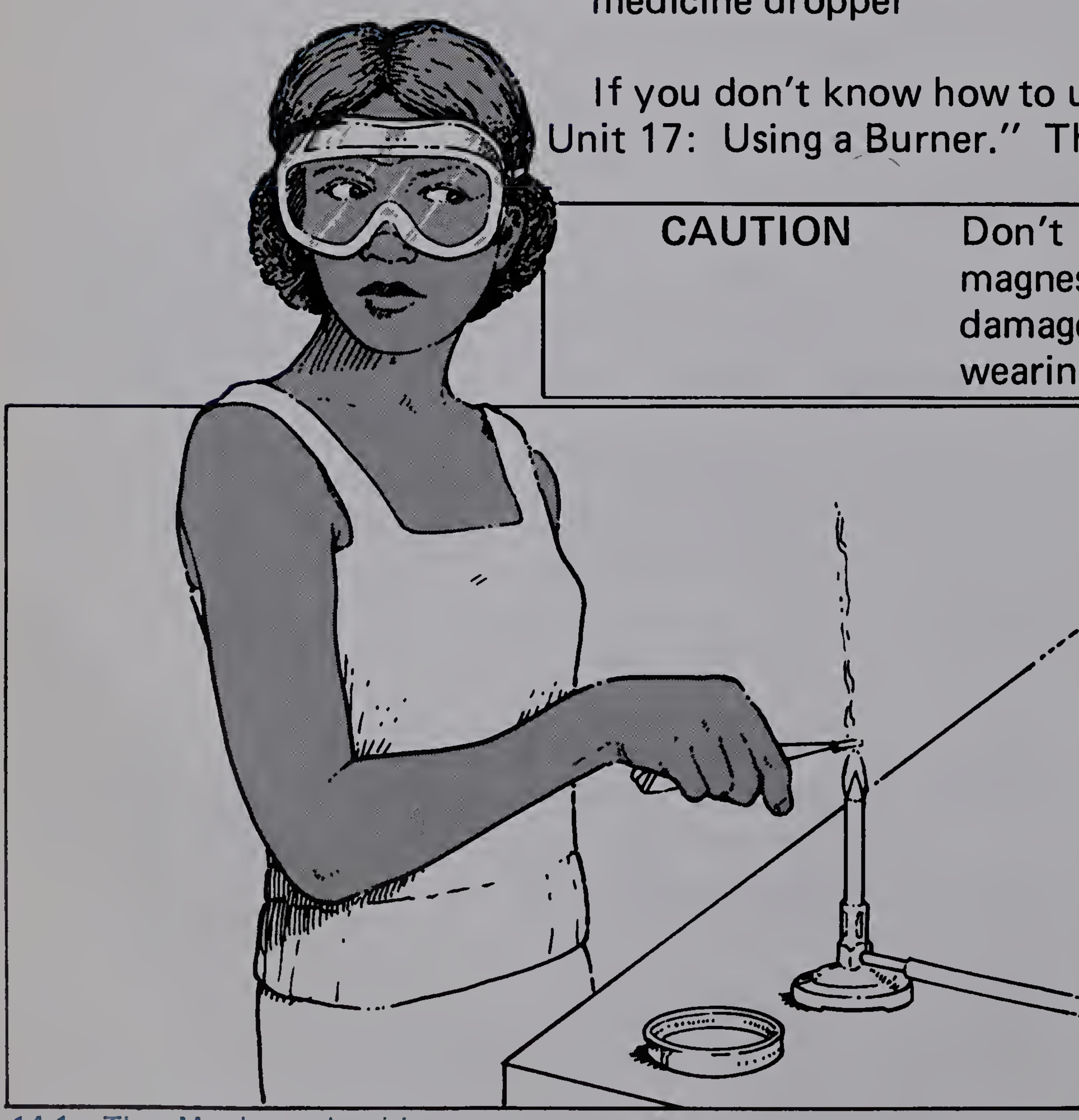
A color chart and pH paper might be used instead of litmus paper.

Students should not touch the hot Mg ribbon or granules. Magnesium burns are often severe and may be slow to heal.

- safety goggles
- 2 g granular magnesium, Mg
- 2-cm strip of magnesium ribbon, Mg
- safety matches
- crucible and cover
- clay triangle
- crucible tongs
- ring stand
- ring clamp
- balance
- tweezers
- steel spatula
- Bunsen burner
- metal can lid
- red litmus paper
- medicine dropper

If you don't know how to use a burner properly, do "Resource Unit 17: Using a Burner." Then begin Step A.

CAUTION	Don't look directly at the burning magnesium, Mg. The light could damage your eyes. Be sure you're wearing your safety goggles.



14-1. The Mg burned with an intense light, forming a white solid.

A. Light the burner, and adjust it to a hot flame. Hold one end of the 2-cm Mg strip with the tweezers.

B. Ignite the Mg, and hold it over the can lid.

- 14-1. Tell what happened and describe the product.

● 14-2. Which gas — nitrogen or oxygen — reacted with the magnesium? What evidence supports your answer?

14-2. Oxygen reacted. A white solid, MgO, rather than a green solid, Mg_3N_2, was formed.

Even though there is much more nitrogen in the air than oxygen, the oxygen reacts more readily.

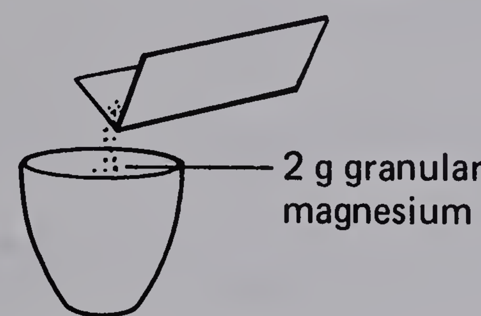
● 14-3. Based on the reaction of air with magnesium, which gas — nitrogen or oxygen — is more reluctant to react?

14-3. Nitrogen, N_2

You can make nitrogen react with magnesium. But to do it, you have to keep some of the oxygen away. One way to do it is to heat the magnesium in a covered dish (a crucible). After the oxygen in the dish is used up, then the nitrogen will react.

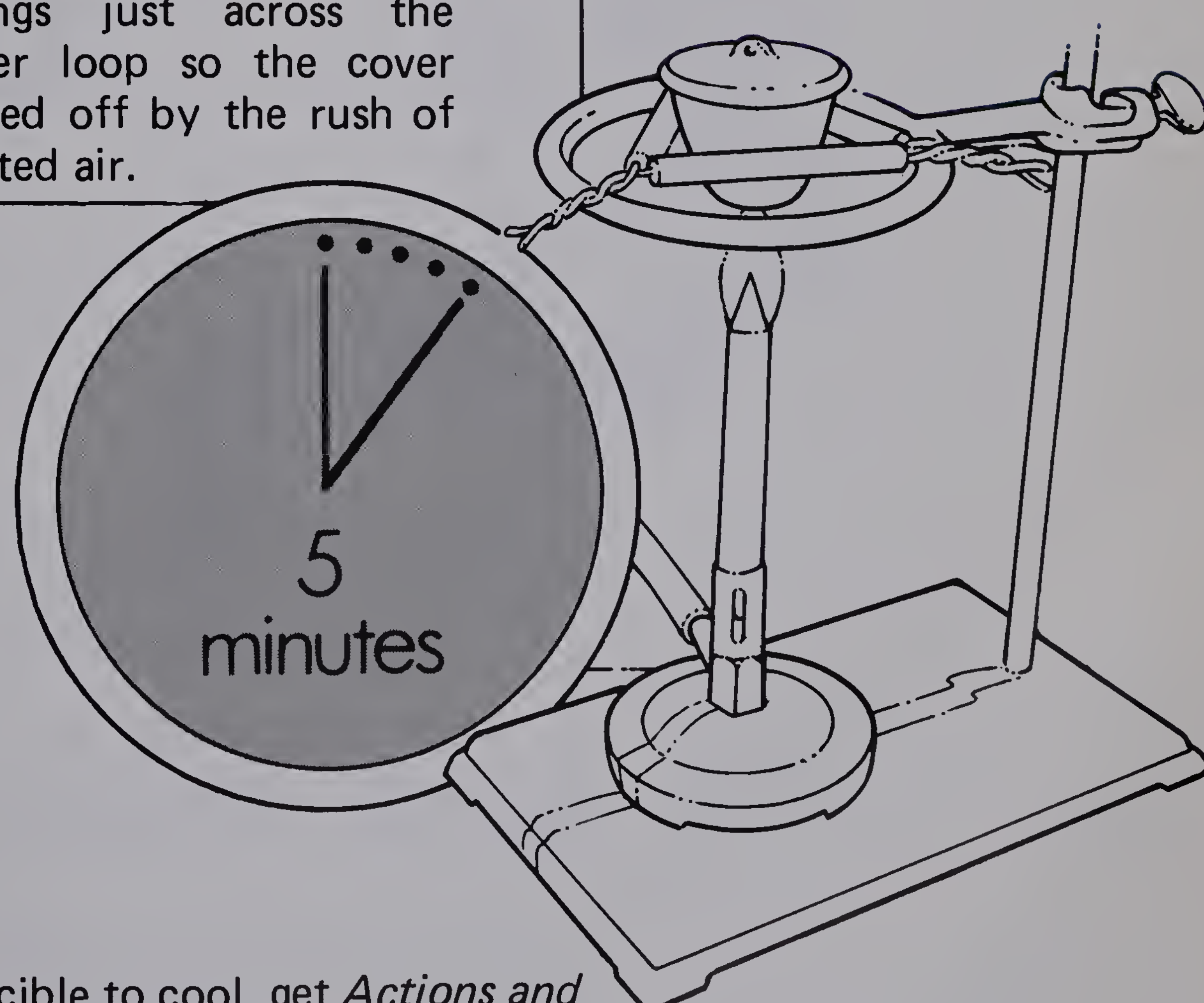
Because of fire potential, Mg granules should be kept away from open flames.

C. Weigh out 2 g of granular magnesium, and put it into the crucible. Place the crucible in the clay triangle on the ring stand. Cover the crucible.

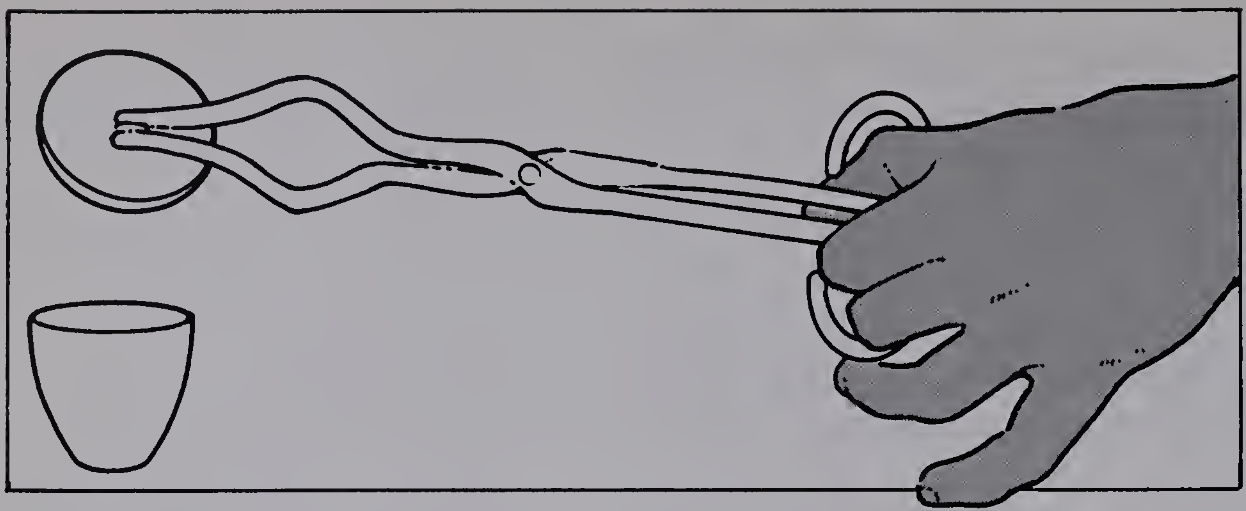


CAUTION	Be prepared for flame and magnesium oxide smoke as magnesium bursts into flame. Hold the crucible tongs just across the crucible cover loop so the cover is not knocked off by the rush of suddenly heated air.

D. Heat the crucible until it glows a dull red. Continue to heat it for another 5 minutes. Then, turn off the burner and let the crucible cool for at least 15 minutes. Don't peek inside the crucible. Re-ignition might occur, and that would ruin the investigation.



While you are waiting for the crucible to cool, get *Actions and Reactions*. Review the activity on chemical bonds if you need to. You'll need that information to be able to explain what happened in your investigation.



E. Using crucible tongs, uncover the crucible and inspect the contents. Move the contents around gently with the steel spatula to see what might be underneath.

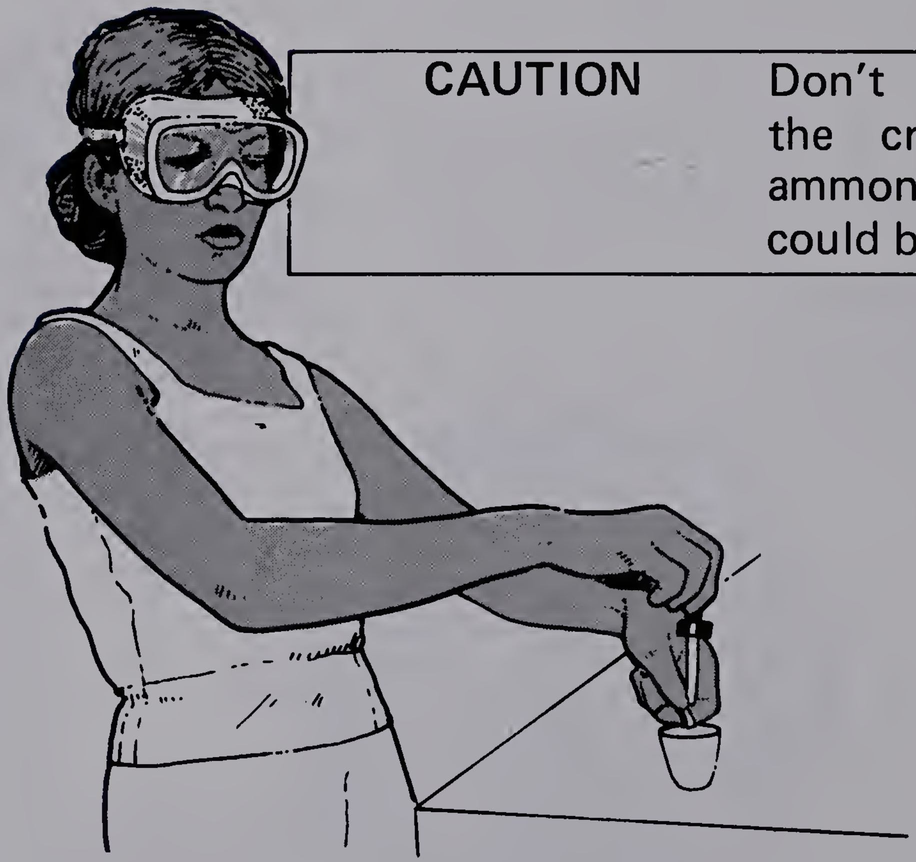
14-4. The crucible contains some white magnesium oxide, some gray black solid, and a small amount of green solid.

14-5. A green solid, Mg_3N_2, formed. [The green solid may not be readily apparent. It is usually at the bottom under the MgO.]

- 14-4. Describe the appearance of the contents of the crucible.

- 14-5. What evidence is there that nitrogen reacted with some of the magnesium?

The appearance of the product may not be very convincing. However, there is another way to detect magnesium nitride besides looking for the green solid. Magnesium nitride reacts with water to form ammonia. You can smell ammonia. It also turns red litmus blue.



CAUTION	Don't get your face too close to the crucible as you sniff for ammonia, NH_3. The NH_3 smell could be very strong.

F. Slightly moisten a piece of red litmus paper. Get a few drops of water in a medicine dropper.

G. Holding the litmus paper above the crucible, add a drop or two of water to the crucible. Carefully sniff for ammonia.

14-6. The litmus paper changed from red to blue. There was an odor of NH_3. The green matter whitened.

- 14-6. What evidence did you get that wetting the contents of the crucible formed ammonia, NH_3? What happened to the green matter in the crucible after it was wet?

As you heated the crucible, some Mg burned to white magnesium oxide, MgO. This used up oxygen, O_2, in the air faster than new air could seep in under the crucible cover. So the remaining heated Mg combined with the nitrogen, N_2, that was in the crucible to form green magnesium nitride, Mg_3N_2.

● 14-7. What evidence did you get from the litmus paper and the odor that there was magnesium nitride in the crucible?

14-7. When water was added, ammonia did form, as evidenced by the odor and the red litmus turning blue.

In your investigation, oxygen reacted very readily with the magnesium. Nitrogen was reluctant to react but did so when most of the oxygen was gone. But there's no point in even trying to get the argon in the air to react. It won't react with anything!

Ar and other inert gases, most notably Kr and Xe, do form complex (coordination) compounds that are unstable. These are unusual compounds.

☆ **14-8. What are the relative reactivities of O_2, N_2, and Ar? What shows that?**

14-8. O_2 reacts more readily than N_2. Ar doesn't react at all. When Mg burns in lots of air, MgO but not Mg_3N_2 forms. Only when there's very little O_2 left will the N_2 react to form Mg_3N_2. The Ar never reacted at all.

You may be wondering by now why there's such a difference in how these three gases react. Chemists explain these differences in chemical reactivity by looking at the structures of each of the three kinds of atoms. Their structures are different, and so their behaviors are different.

The chemical reactivity of the gases can be traced back to the number and arrangement of the outermost electrons of their atoms. Look at Figure 14-1 below.

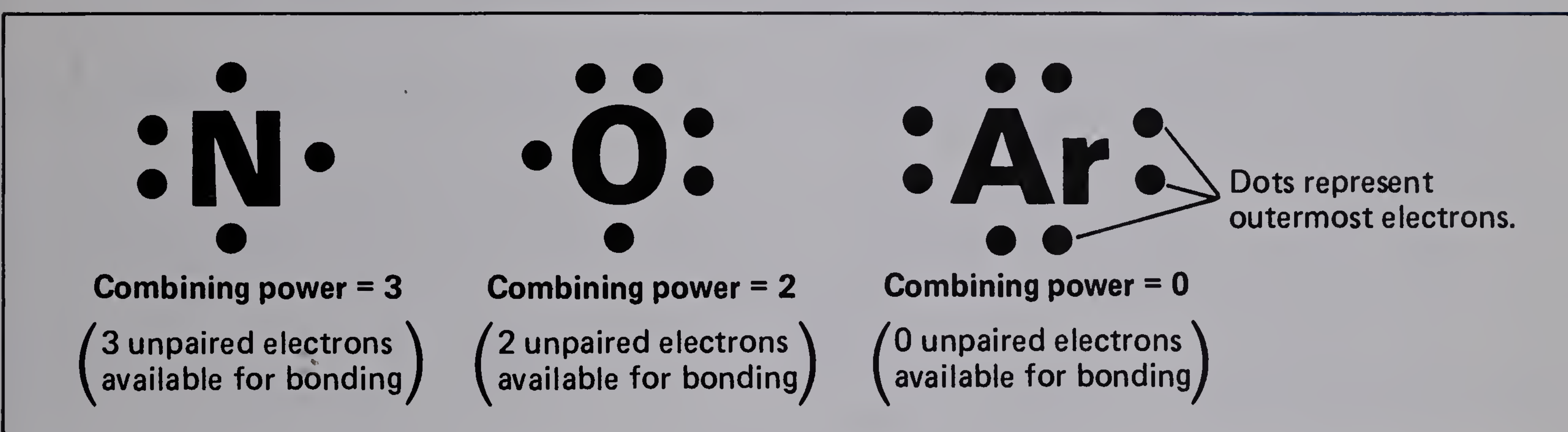


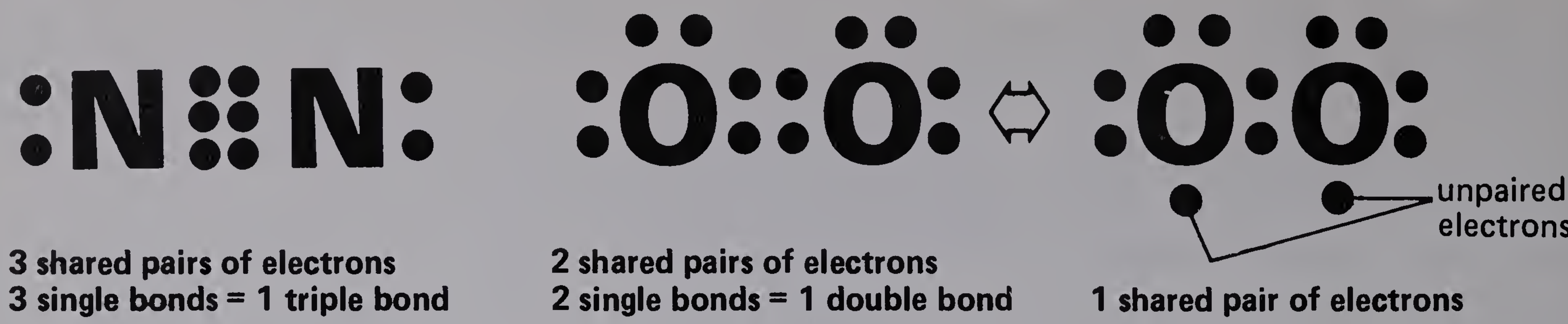
Figure 14-1

As you can see in Figure 14-1 above, combining power is an indicator of how many unpaired outermost electrons an atom has. Thus, it is also an indicator of how many bonds an atom will form.

● 14-9. If each unpaired electron of an atom can form a bond, how many bonds can a nitrogen atom, N, form? An oxygen atom, O? An argon atom, Ar?

14-9. N, 3 bonds; O, 2 bonds; Ar, no bonds

According to Figure 14-1 above, nitrogen atoms have three unpaired electrons. To pair these up, two nitrogen atoms come together and form a nitrogen molecule with three shared pairs of electrons — a triple bond. See Figure 14-2 on page 68.



N_2 plays fair and square, so you have to break three bonds to get it apart into atoms.

Oxygen cheats a little and switches from two bonds to a single bond. You don't have to break two bonds to get it apart into atoms. And having unpaired electrons makes it much more reactive than N_2.

Figure 14-2

The electron-dot model is inadequate for molecular oxygen. The atoms are bound with a more complex bond than a single bond. However, a double bond is inaccurate and also fails to show the unpaired electrons. The only satisfactory representation of the bond involves molecular orbitals. You may wish to develop this with abler students; however, the explanation given here is adequate for the objective of this activity.

● 14-10. Why is N_2 harder to separate into atoms than O_2?

14-10. N_2 molecules are more strongly tied together — by triple bonds — than O_2 molecules — by single and double bonds.

The nitrogen in the air is in the form of diatomic (two-atom) molecules. The triple bond (three single bonds) in these molecules is quite strong. And for nitrogen to react with other substances, such as magnesium, all three bonds must be broken. This takes a lot of energy — heat energy, for example.

● 14-11. Why was it hard to get nitrogen gas to react with magnesium?

14-11. All three bonds in the N_2 molecules had to be broken.

The oxygen in the air is in the form of diatomic molecules too. But the bond between the oxygen atoms is a weaker bond than that in the N_2 molecule. It is easier to free the oxygen atoms to react with other substances. Oxygen molecules also have unpaired electrons. This too makes them more reactive than nitrogen molecules.

Oxygen reacts readily. Burning reactions of all kinds involve oxygen — even the "burning" of food in your body to produce energy. You breathe in air, and the oxygen in it reacts in your body. But the nitrogen doesn't — you breathe the N_2 right out again. N_2 is not reactive enough to do anything in your body.

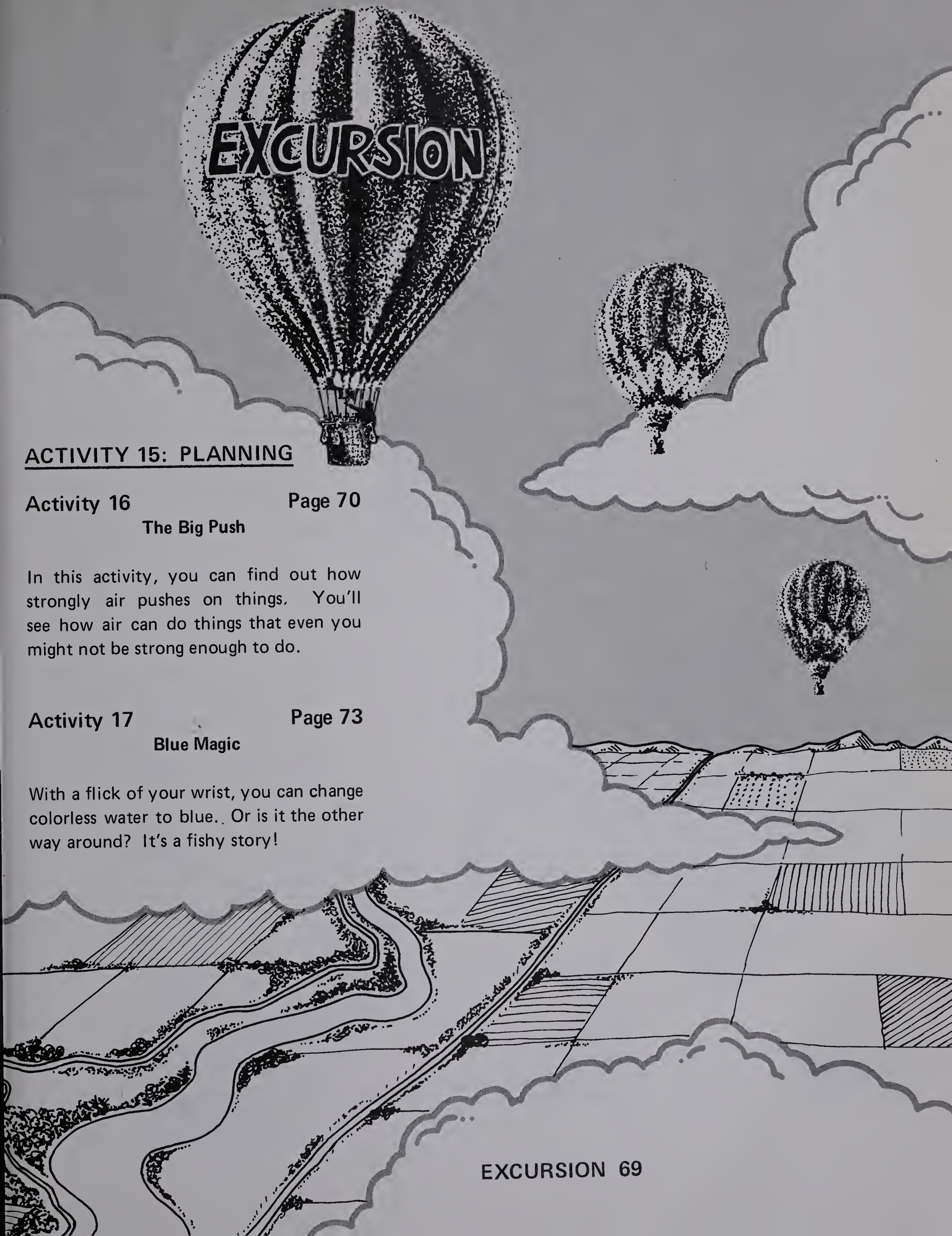
● 14-12. Explain how the structure of O_2 molecules accounts for the fact that they are so much more reactive than N_2 molecules.

14-12. O_2 has only one bond and has two unpaired electrons, whereas N_2 has three bonds and no unpaired electrons.

Argon, Ar, is completely unreactive — it doesn't even form diatomic molecules with itself! The atom has no unpaired electrons and no tendency to bond, so its combining power is zero.

☆ **14-13. Explain the difference in reactivities of nitrogen, oxygen, and argon gases in terms of their structures.**

14-13. The oxygen atoms in the diatomic molecule are not strongly held together, and each contains an unpaired electron. Nitrogen atoms are strongly held together in the diatomic molecule with a triple bond. Argon has no unpaired electrons and no tendency to react.



ACTIVITY 15: PLANNING

In this activity, you can find out how strongly air pushes on things. You'll see how air can do things that even you might not be strong enough to do.

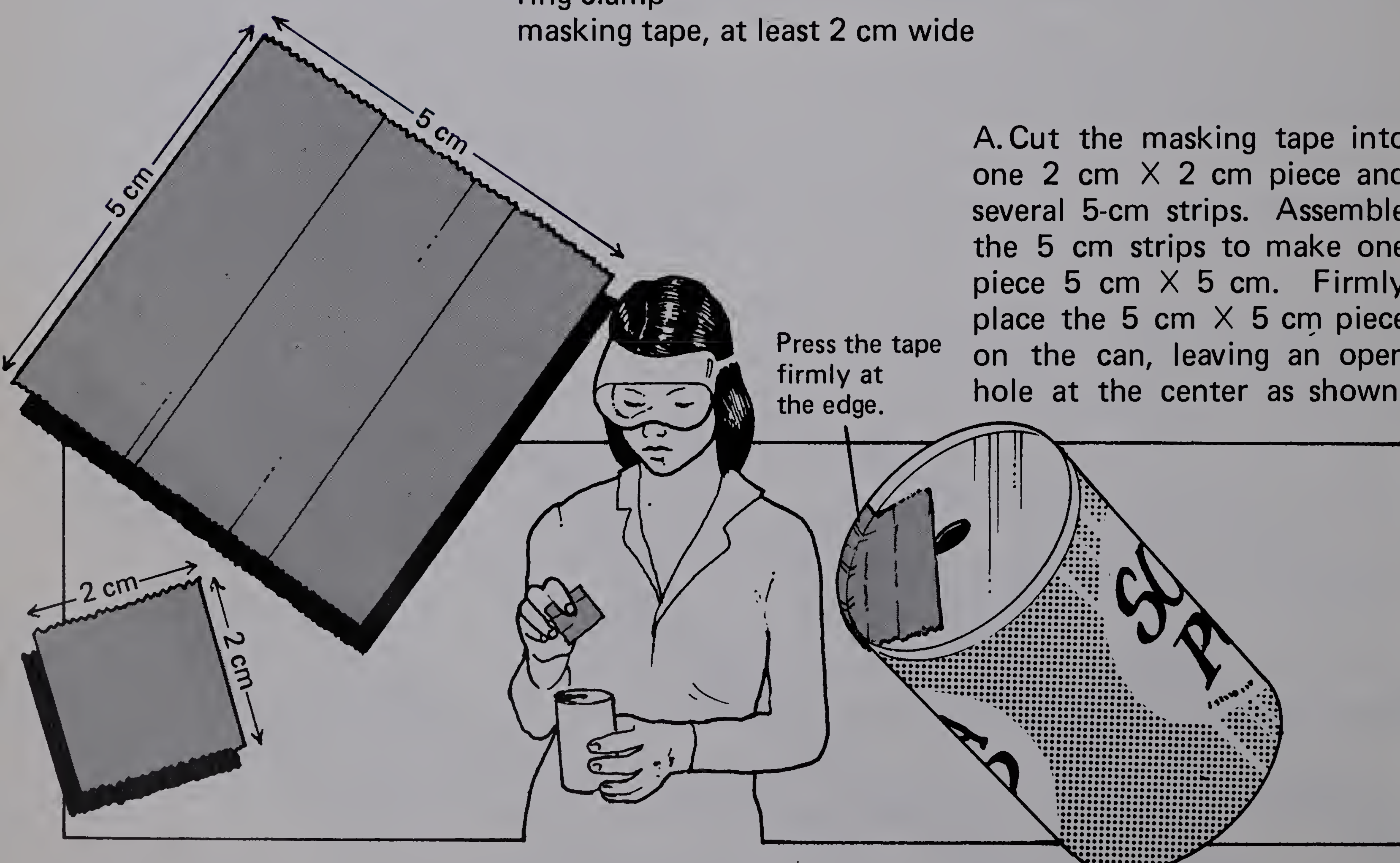
With a flick of your wrist, you can change colorless water to blue. Or is it the other way around? It's a fishy story!

ACTIVITY 16: THE BIG PUSH

Earth's gravity pulls on air just as it pulls on things you can see. In fact, air is so heavy that you carry around almost seven hundred kilograms (about two-thirds of a tonne) of air on your head and shoulders! What keeps you from being crushed by all that air is your inside pressure. It pushes back on the outside air pressure.

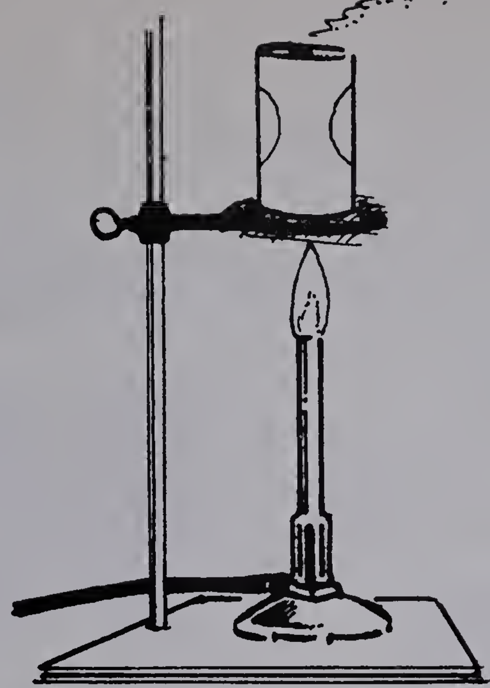
You probably need to see that to believe it. And you can — in a way. See what would happen if you let the air push on something without anything on the other side to push back. You will need these materials.

- safety goggles
- beaker tongs
- scissors
- Bunsen burner or other heat source
- 50-ml graduated cylinder
- wire gauze
- empty, thin aluminum beverage can
- heavy glove or mitt-type pot holder
- ring stand
- ring clamp
- masking tape, at least 2 cm wide



A. Cut the masking tape into one 2 cm X 2 cm piece and several 5-cm strips. Assemble the 5 cm strips to make one piece 5 cm X 5 cm. Firmly place the 5 cm X 5 cm piece on the can, leaving an open hole at the center as shown.

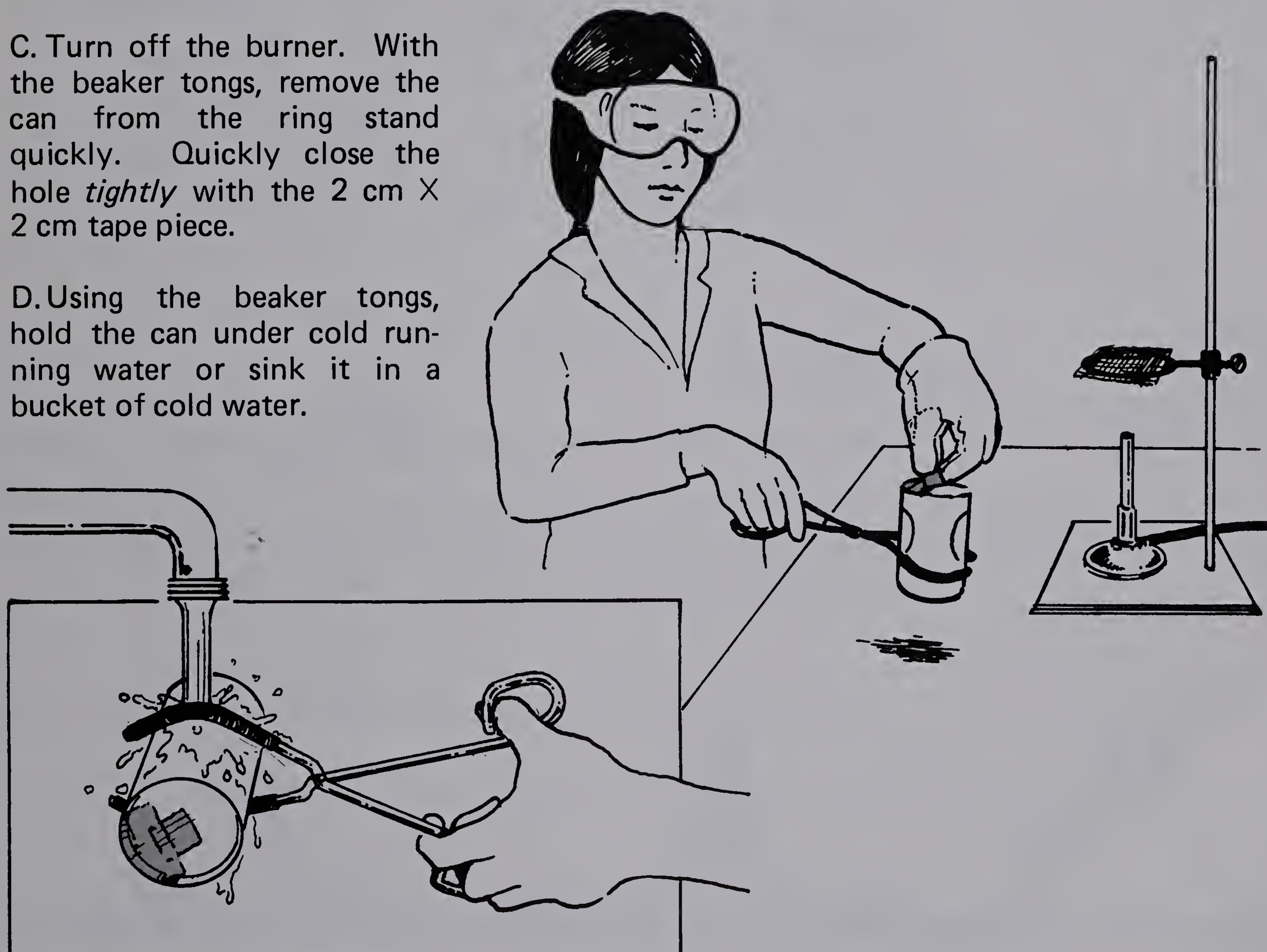
B. Put about 10 ml of water into the can. Put the can on the wire gauze on the ring stand and heat it. Boil the water for one minute or so after the steam appears. (You'll hear it boil.)



CAUTION	Don't touch the hot can with bare hands.

C. Turn off the burner. With the beaker tongs, remove the can from the ring stand quickly. Quickly close the hole *tightly* with the 2 cm X 2 cm tape piece.

D. Using the beaker tongs, hold the can under cold running water or sink it in a bucket of cold water.



- 16-1. What happened to the can?

- 16-2. How did boiling the water help get rid of the air inside the can?

16-1. It collapsed.

16-2. When the water boiled, it changed to a gas and pushed out the air.

Figure 16-1 below explains what happened to the can.

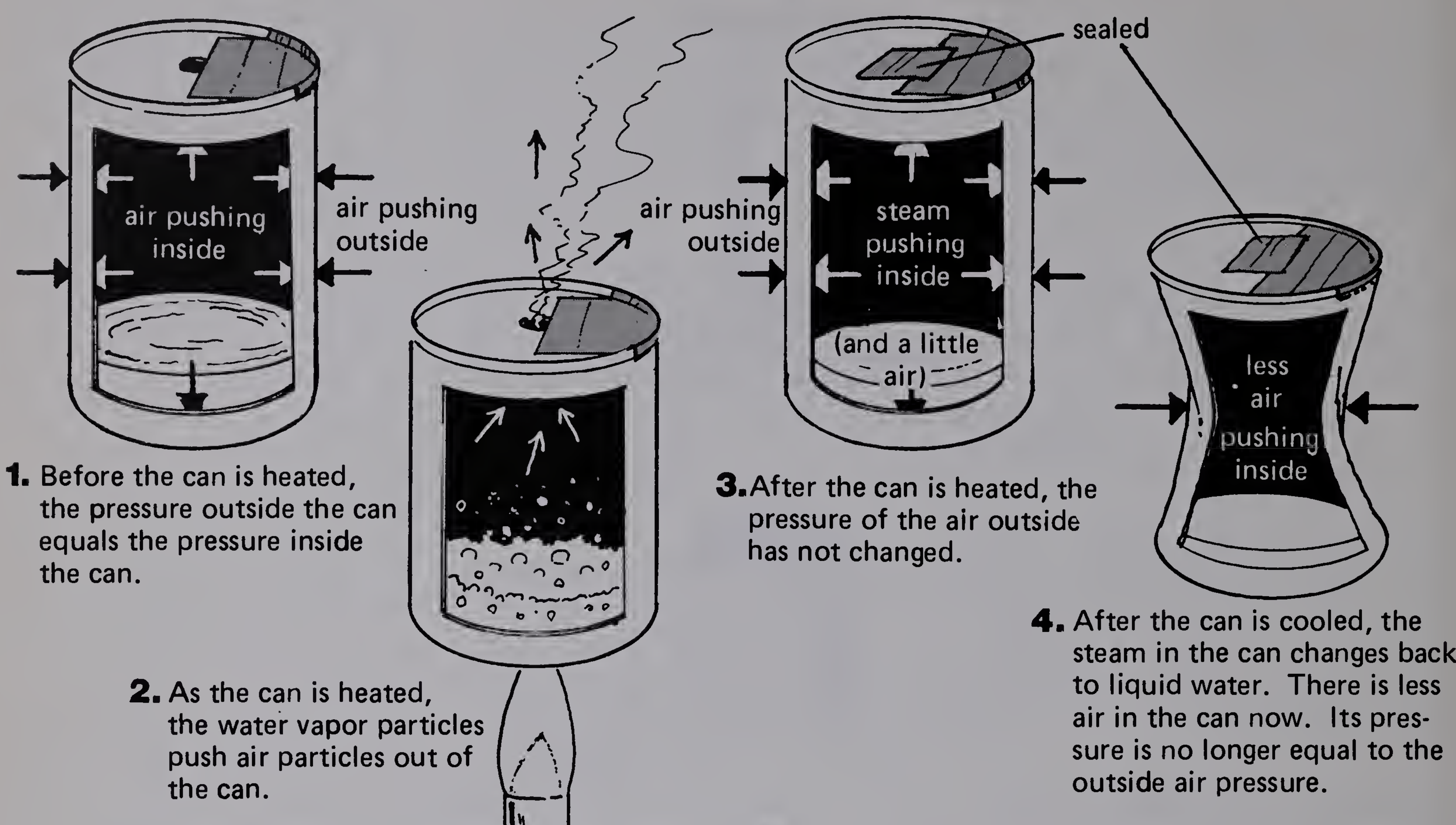


Figure 16-1

16-3. More air pressure is pushing on the outside of the can than is inside pushing back.

16-4. Air has been forced out of the suction cup. With less air inside, there is more pressure outside.

☆ **16-3. Why does a sealed, steam-filled can collapse as it cools?**

A suction cup is based on the same idea. The cup seems to be holding or sticking to the surface. But actually it's being pushed against the surface by the outside air. Look at Figure 16-2 below.

Figure 16-2

16-5. You let air in under the cup so that the pressure is equal on both the inside and the outside of the cup.

☆ **16-4. How does a suction cup hold to a surface?**

• 16-5. How do you get a suction cup to release?

ACTIVITY 17: BLUE MAGIC

ACTIVITY EMPHASIS: A simulation of the addition and removal of oxygen in river water is made by shaking a bottle containing an "oxygen-demanding organic" — dextrose. The shaking oxygenates the water. Then, the dextrose combines with the oxygen. Methylene blue indicator is used to signal these changes.

MATERIALS PER STUDENT LAB GROUP: See tables in "Materials and Equipment" in ATE front matter. See "Advance Preparations" in ATE front matter.

Animals and plants that live on land or in the air can get all of the oxygen they need from the air. They do so by respiration. But there are lots of animals and plants that live in water. Where does their oxygen come from?

It comes from the air. And it comes from photosynthesis by water plants and microorganisms such as the green algae. The air that dissolves in water provides dissolved oxygen for use by animals and plants that live in water. Photosynthesis by plants also produces oxygen that dissolves in the water. But the oxygen dissolved in water gets used up. Figure 17-1 below shows some users of oxygen.

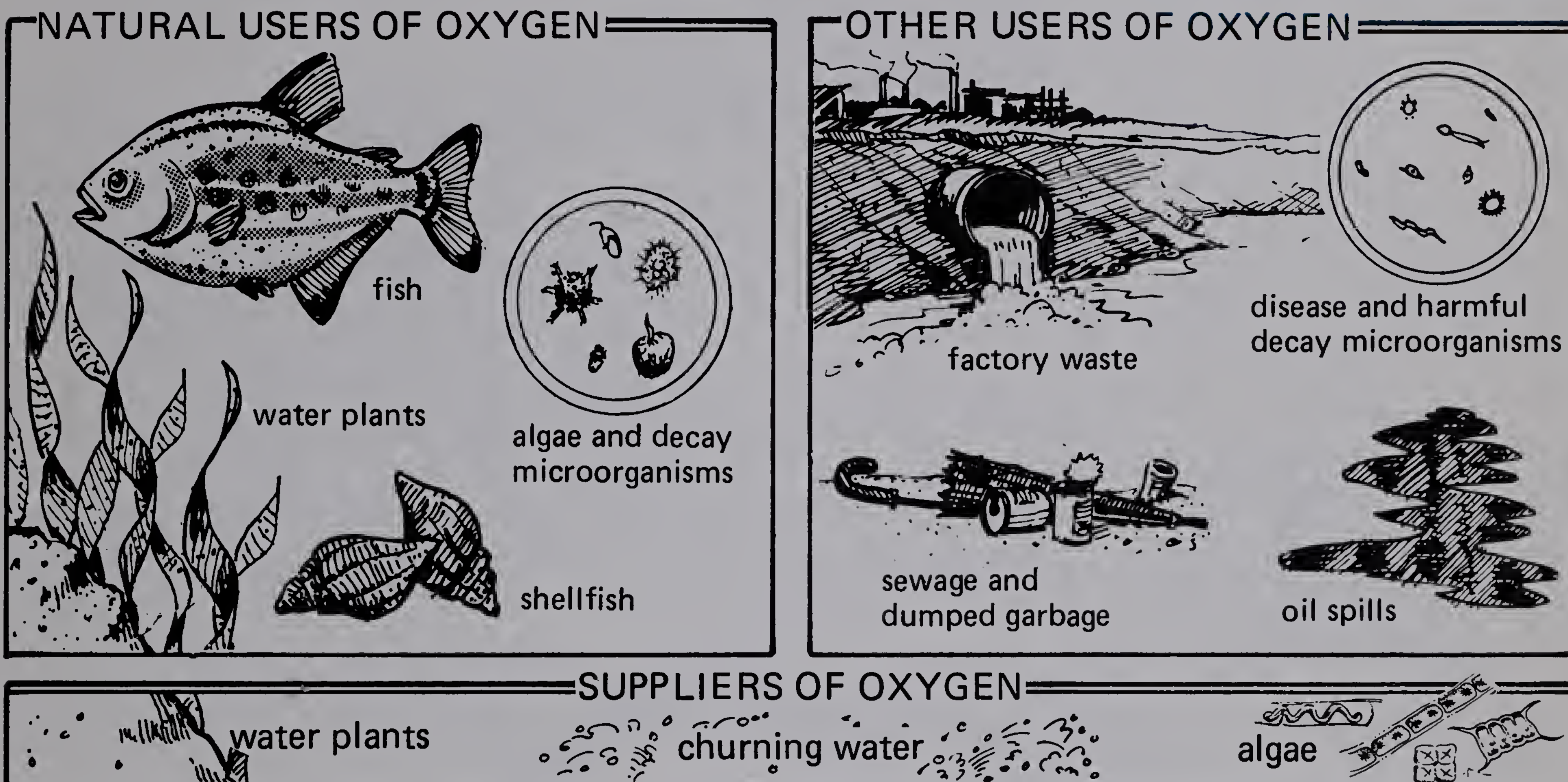


Figure 17-1

☆ **17-1. What kinds of things put oxygen back into the water?**

17-1. Churning water, water plants, and algae

● 17-2. Which living things use oxygen but don't put oxygen back into the water?

17-2. Animals, such as fish and shellfish

● 17-3. What do people put into water that might take oxygen away from living things in the water?

17-3. Sewage and dumped garbage, factory wastes, spilled oil, and disease and harmful decay microorganisms

The oxygen that is used up has to be put back if plants and animals are to keep living in the water. The green plants put some oxygen back. And some is put back by the dissolving of oxygen from the air in water.

17-4. From green plants and by dissolving from the air

☆ **17-4. What are two ways in which oxygen can be put back into water?**

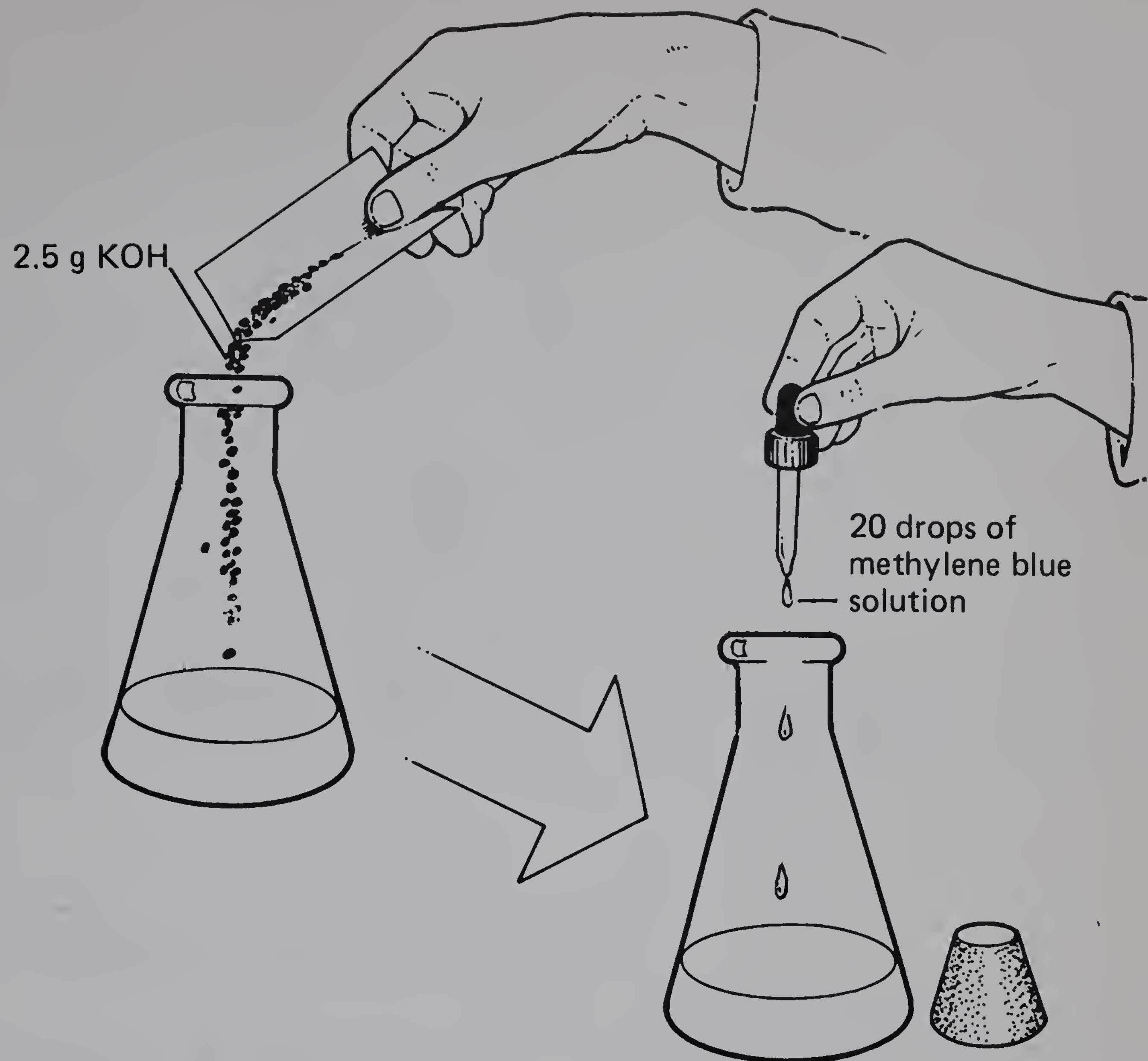
There would be more oxygen for fish and shellfish if there was less pollution of water by wastes and sewage. There would also be more oxygen if there were more plants in water and more dissolving of oxygen by water. How can more oxygen be dissolved in water to replace the oxygen that is being used up? Let's see. You'll need the following materials.

safety goggles
100-ml graduated cylinder
500-ml Erlenmeyer flask
stopper to fit flask
1.5 g dextrose
2.5 g potassium hydroxide, KOH
methylene blue solution in dropping bottle
balance

CAUTION	Potassium hydroxide is very corrosive. Be careful not to get any on your hands or clothing. In case any spills, wash with plenty of water. Tell your teacher if any spills.

A. Measure 125 ml of water into the Erlenmeyer flask. Weigh out 1.5 g of dextrose, and add it to the flask. Swirl the flask until the solid completely dissolves.

B. Weigh out 2.5 g of potassium hydroxide, KOH. Be sure to use weighing paper. Add the potassium hydroxide to the flask. Again swirl the flask until the solid dissolves.

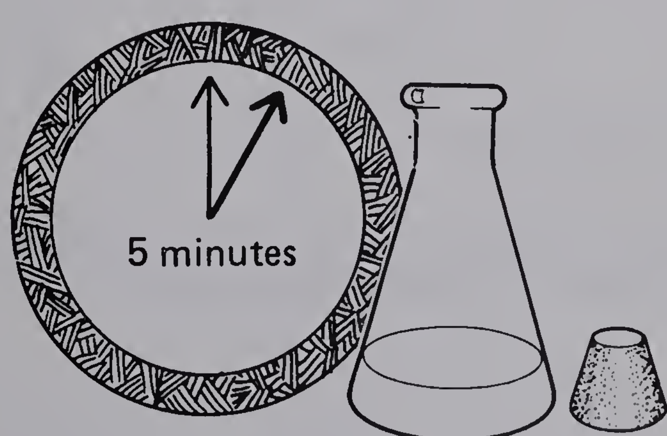


C. Add 20 drops of methylene blue solution to the flask. Again, swirl the flask until the blue color is even throughout the liquid.

When there is oxygen dissolved in water, methylene blue is blue. When there is no oxygen dissolved, methylene blue turns colorless. In the presence of potassium hydroxide, dextrose uses up oxygen as if it were one of the oxygen users shown in Figure 17-1 (page 73).

Methylene blue is colorless until it combines weakly with dissolved oxygen. Then it turns blue. When the oxygen reacts with dextrose, the methylene blue, no longer combined with oxygen, becomes colorless.

D. Set the flask aside. Don't touch it for at least 5 minutes.



• 17-5. What happened to the color of the water as you let the water sit for five minutes?

17-5. The blue disappeared.

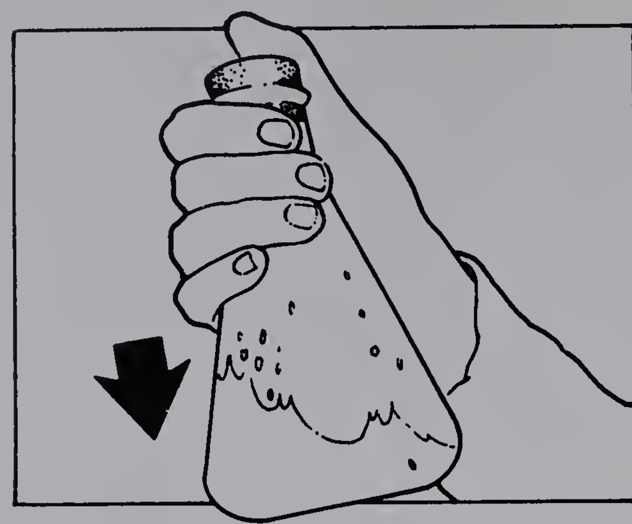
• 17-6. What must have happened to the dissolved oxygen as the water sat?

17-6. The dissolved oxygen must have been used up.

• 17-7. What was in the water that used up the dissolved oxygen?

17-7. Dextrose

One way to get more oxygen to dissolve in the water is to mix the water with the air above it.



E. When the blue color has disappeared, stopper the flask. Put your finger on the stopper, as shown. Give the flask one hard downward jerk.

17-8. The solution turned blue again, indicating oxygen is dissolved in the water again.

● 17-8. What happened to the color of the solution when it was shaken? What does that tell you about the solution?

F. Repeat Steps D and E as many times as you want to.

Look at Figure 17-2 below. It reviews what happened in the flask.

Observant students may notice a blue film remaining on the liquid surface as the liquid turns colorless — a result of O_2 entering the water at the surface.

After an hour or so, the methylene blue becomes ineffective as an indicator here. The solution will become yellowish.

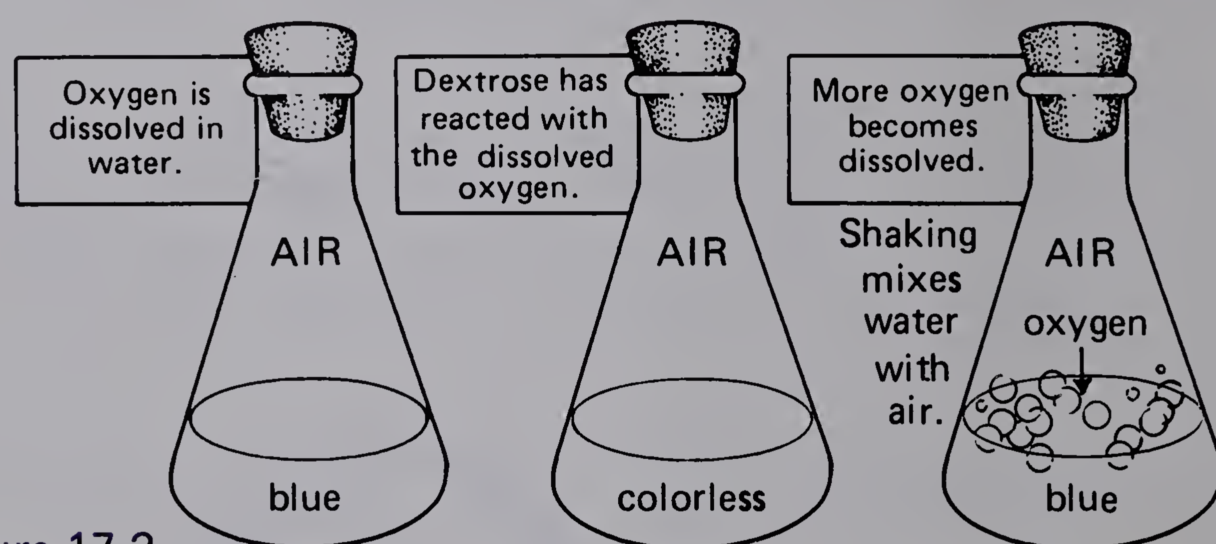


Figure 17-2

17-9. More oxygen will dissolve in churning water. Shaking lets more water come in touch with the oxygen in the air.

☆ **17-9. Will more oxygen dissolve in churning water or in still water? Explain your answer.**

Tumbling streams are richer in oxygen than quiet waters. A pool or lake fed by such streams tends to have more oxygen and more fish in it than water fed by quiet, slow-moving streams. That's the reason some people catch more fish — they know where the oxygen is!

Waves breaking in shallow water or on beaches and shores provide more oxygen than wave motion does to deep seas farther out. That's the reason sand bars, reefs, and the sea bottom close to shore are full of life. But the deep ocean is a watery desert where life is harder and rarer.

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PERSONNEL

The ISIS Project is an intricate effort involving many people in many roles. The following individuals have made significant contributions to that effort.

Project Staff

Ernest Burkman, Director
William R. Snyder, Associate Director

Tedd Arnold
Gary K. Baker
Joe Beditz
Denis Blakeway
Calvin E. Bolin
Drennen A. Browne
Robert Buchanan
Marcia Bujold
Jack J. Bulloff
David L. Camp
Gwendie Camp
Ace Carroll
Maxwell Caskie
Robert Celander
Richard C. Chittenden
Jerome L. Ciesla
Clifton Bob Clark
Robert L. Cocanougher
Sara P. Craig
Julia Damon
Stewart P. Darrow
Allan D. Dawson
Joel Dawson
Cheval Fagan
Gene Floersch
Ronald N. Giese
Gail M. Grandy
James A. Greenwood
James P. Hale
Fred Hartford
James A. Hathway
Lila T. Kirschbaum
Jeanne C. Klinzing
John R. Knudstrup
Michael Krone
Ronald C. Laugen
Francis X. Lawlor
Clarke G. Lieffers
Pam Little
Robert C. Loser
Adrian D. Lovell
Elisabeth McCurnin
Dawn McQueen
Bernadette R. Menhusen
Brenda Musgrave-Propst
Gerald Neufeld
Hugh Nicholson
Betty Jerrido Oates
Barney Parker
Marvin D. Patterson
Charles E. Peters
Ronald J. Polland
David D. Redfield
Susan Reichman
George A. Reid, Jr.
Jenne Taylor Richardson
John Roberge
Dee Dee Shand
Beverly Smith
Donald A. Smith
John A. Sumner
Clifford Swartz
Mike Tillmans
Ralph G. Vedros
Robert Vickery
Brian Webber
Stuart H. Weinstein
Lois S. Wilson
Jay A. Young
Brenda Zeh

Writing Conference Participants and Author-Consultants

Betsy Balzano, SUNY at Brockport; David A. Bare, Florida State University; David Berey, Roslyn (NY) Schools; Robert Bernoff, Penn. State University; Capt. George Bond, Naval Coastal Systems Laboratory; Ted Bredderman, Delmar, New York; John Cunningham, Keene State College; James DeRose, Marple-Newton (PA) School District; Robert D. Eddy, Tufts University; I. Dwaine Eubanks, Oklahoma State University; Roy Gallant, Rangeley, Maine; Orrin Gould, University of Illinois; Francis U. Hackley, Leon (FL) Schools; Robert C. Harriss, Florida State University; Jack Hassard, Georgia State University; Robert E. Horvat, SUNY at Buffalo; Stuart J. Inglis, Medford (OR) School District; Jane Kahle, Purdue University; Al Kaskel, Evanston (IL) Schools; David Klasson, Fall River (CA) Joint Unified School District; David Kuhn, Tarrytown (NY) Schools; Clarence T. Lange, Clayton (MO) Schools; Sander Latts, University of Minnesota; Ray D. Lauer, Florida State University; Robert L. Lehrman, Roslyn (NY) Schools; Harleen McAda, University of California at Santa Barbara; Wendell Mohling, Shawnee Mission (KS) Schools; Floyd Monaghan, Michigan State University; Terrence G. Oas, Florida State University; Rod O'Connor, Texas A & M University; M. Larry Peck, Texas A & M University; William S. Ravenel, Florida State University; Shirley Richardson, San Diego (CA) Schools; Guenter Schwarz (Deceased); John S. Shelton, La Jolla, California; Douglas P. Smith, Florida State University; Timothy B. Vanderwood, Florida State University; Claude A. Welch, Macalester College, Joe Dugan Whiteside, Florida State University; Owen York, Kenyon College

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